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CHEMICAL NOTES FOR THE LECTURE-ROOM.

SECOND SERIES.

PND

LONDON: PRINTED BY
SPOTTISWOODE AND CO., NEW-STREET SQUARE
AND PARLIAMENT STREET

NOTES ON THE METALS:

BEING

A SECOND SERIES

OF

CHEMICAL NOTES FOR THE LECTURE-ROOM.

By THOMAS WOOD, Ph.D., F.C.S.

LECTURER ON CHEMISTRY AT THE BRIGHTON COLLEGE
ETC.

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LONDON:
LONGMANS, GREEN, AND CO.
1868.

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ASTOR, LENOX AND
TILDEN FOUNDATIONS.
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PREFACE.

THE AUTHOR has long felt in the course of his duties as Lecturer on Chemistry the anomalous position of a master who had to employ for something important in the science an expression to which he could not assign any definite meaning. This has been the case with the term 'acid.'

As all manuals in general use gave the word acid freely and promiscuously, it was found impracticable to do away with the term while these books sanctioned its constant application.

The requirement of a small book of 'notes on the metals,' for the pupils at Brighton College, has been taken advantage of to ascertain how far the writer's views are correct. They are these: 1. The term acid is not necessary in teaching chemistry. 2. Clearer and more accurate views on chemical subjects will be instilled into the minds of beginners without it than by using it as is done at present in most class-books and classes. 3. Its employment, even in the most limited degree, without assigning to it a definite meaning, is improper, since it leaves only a hazy idea of what is intended, where clearness is one of the first objects to be, if possible, attained.

The term 'acid' is therefore not applied in these notes.

Some information, it may be thought, might have been omitted, and much more of what is considered essential substituted with advantage. The work, however, has already exceeded the limits originally intended, and indulgence is claimed for the very great difficulty of expressing new views in a simple and condensed form.

The references in the text (see *C. N.* p.   ) are to the second edition of the Author's 'Chemical Notes.'

T. W.

LABORATORY: 23 GROVE STREET, LISSON GROVE, N.W.
September 29, 1868.

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NOTES

ON

METALS.

PHYSICAL PROPERTIES OF THE METALS.

THE METALS are far more numerous than the non-metals, being about 50 in number. (See p. 2.) Although they differ widely from one another in their chemical and physical properties, yet they possess certain peculiarities in common, both with each other and with the non-metals, rendering it extremely difficult to make any clear and satisfactory classification of them. (See *Chemical Notes*, p. 24, second edition.)

The following are the principal characteristic physical properties belonging to the metals:—

1. METALLIC LUSTRE.—All metallic substances, after they have been melted and polished, possess in a high degree that peculiar property of reflecting light, called *metallic lustre*. Some non-metallic bodies, as iodine, graphite, &c. also possess considerable lustre, but not in so marked a degree as the metals.

2. COLOUR.—There is a great similarity in the colour reflected by the different metals, which, with few exceptions, varies only between the white of silver, the grey of iron, and the bluish colour of zinc and lead. The exceptions are: copper, which is the only red metal; gold, which

LIST OF THE METALS.

Common Metals.	Symbols.	Combining weight.	S. g.
Lithium	Li	7	0·59
Sodium (*Natrium*)	Na	23	0·97
Magnesium	Mg	24	1·74
Aluminium	Al	27	2·6
Potassium (*Kalium*)	K	39	0·86
Calcium	Ca	40	1·58
Chromium	Cr	52·5	6·8
Palladium	Pd	53·2	11·4
Manganese	Mn	55	8·00
Cadmium	Cd	55·7	8·6
Iron (*Ferrum*)	Fe	56	7·84
Cobalt	Co	59	8·95
Nickel	Ni	59	8·80
Copper (*Cuprum*)	Cu	63·5	8·95
Zinc	Zn	65	7·00
Arsenic	As	75	5·80
Strontium	Sr	87·5	2·54
Silver (*Argentum*)	Ag	108	10·50
Tin (*Stannum*)	Sn	118	7·30
Antimony (*Stibium*)	Sb	122	6·70
Barium	Ba	137	4·00
Gold (*Aurum*)	Au	197	19·30
Platinum	Pt	198	21·50
Mercury (*Hydrargyrum*)	Hg	200	13·60
Lead (*Plumbum*)	Pb	207	11·40
Bismuth	Bi	210	9·80

Rare Metals.	Symbols.	Combining weight.	S. g.
Glucinum or Beryllium	G Be	7	2·1
Yttrium			
Erbium			
Terbium			
Norium			
Titanium	Ti	25	
Zirconium	Zr	33·6	
Molybdenum	Mo	46	8·6
Cerium	Ce	47	
Lanthanum	La	47	
Didymium	D	48	
Ruthenium	Ru	51·7	11·4
Rhodium	R	52·2	12
Thorinum	Th	59·5	
Uranium	U	60	18·4
Vanadium	V	68·5	
Rubidium	Rb	85	1·52
Tungsten or Wolfram	W	92	17·6
Niobium or Columbium	Nb	94	
Iridium	Ir	98·7	21·1
Osmium	Os	99·5	21·4
Tellurium	Te	129	6·65
Cæsium	Cs	133	
Tantalum	Ta	172	
Thallium	Tl	204	11·9

is bright yellow; and calcium and barium, which are yellowish.

Many metals, when in a fine state of division, lose the power of reflecting light, and appear as *dark powders*, which, however, on being strongly pressed together and burnished, resume their characteristic lustre.

3. Opacity (from *opacitas*, darkness).—The metals are remarkable for their power of preventing the passage of light, since the thinnest leaves of metal entirely cut off the rays of light, and thus are classed amongst the most opaque substances. There is, however, one exception to the above rule in the metal gold, which, in very thin leaves of not more than the $\frac{1}{200000}$ of an inch thick, allows green light to pass through it.

4. Hardness.—There is a very wide difference in the hardness of the several metals, and in the same metal under altered circumstances. Thus most of the metals are *solids* at the ordinary temperature of the air; one metal, however, is a liquid, namely, mercury or quicksilver (Hg); and could H be condensed to the liquid or solid form, it would most likely, judging from its chemical properties, appear as a metal, thus adding a gaseous metal to the list. Again, iron may be made so soft as to be sawn with ease by a steel saw, or, as steel, so hard as to scratch glass. Potassium is a metal so soft that it can be kneaded between the fingers like putty, and lead may be cut with a pocket-knife.

Elasticity and *sonorousness* are properties belonging to many of the harder metals and their compounds. The well known elasticity of steel renders it peculiarly applicable for the construction of springs; and the ring of good coin is proverbial.

5. Malleability (from *malleus*, a hammer) is the capability of being spread out, and forming thin leaves or sheets, when hammered or rolled; *e.g.* gold is the most *malleable* of all the metals, for it can be beaten out

with a hammer until a square foot weighs less than three grains.

6. BRITTLENESS is the opposite to *malleability*, as it is the property of breaking into smaller pieces when hammered; *e.g.* the most brittle metals, antimony and arsenic, may be powdered in a porcelain mortar. Zinc is a metal which stands midway between the brittle and malleable metals; for when cold it is brittle, but when slightly heated may be rolled into sheets, and as sheet-zinc is used for roofing houses.

7. DUCTILITY (from *duco*, I draw, I lead) is the capability of being drawn out into wire. It is intimately connected with the property of malleability; but gold, although the most malleable metal, is not the most ductile, for platinum is quite equal to it in this respect. A gold wire and a platinum one have both been drawn to the $\frac{1}{30000}$ of an inch in thickness.

8. TENACITY (from *tenax*, holding fast) is intimately connected with ductility, and is the power of a wire to support a weight or withstand a strain. It is found that wires of cobalt and iron are the most tenacious, and that lead is amongst the least, iron being 26 times as tenacious as lead.

The properties of hardness, malleability, brittleness, ductility, and tenacity are greatly dependent upon the processes of *annealing* (from *on*, and *ælan*, to burn)—that is, heating a substance to dull redness, and allowing it to cool slowly, thereby rendering it less brittle; and *pressure*, either between rollers or by hammering, which greatly affects these properties.

9. SPECIFIC GRAVITY or DENSITY (s.g.).—A remarkable difference exists between the comparative weights of the several metals. The liquid called mineral-naphtha, in which some of the metals are kept to protect them from the action of the atmosphere, is lighter than water, and consequently *swims* on it. Lithium, the lightest metal known,

swims when placed in such naphtha. The metal potassium sinks in naphtha, but will swim on the surface of water. Iron, as is well known, will sink in water, but swims if thrown upon mercury; while platinum is nearly twice as dense as mercury.

The s.g. of water is always taken as 1·000. Then the s.g. of lithium is found to be 0·590—that is, a piece of lithium is only rather more than half the weight of a piece of water its own size. The metal platinum has a s.g. of 21·500—that is, it weighs twenty-one and a half times as much as a piece of water of its own size.

10. FUSIBILITY (from *fundo*, I pour out), or the quality of being converted from the solid to the liquid state by the application of heat, is possessed by all the metals. The melting point, or point of fusion, however, of the several metals varies greatly. Thus mercury is liquid much below the freezing point of water. Potassium melts below the boiling point of water. Cast iron cannot be melted in the fire of a smith's forge; and, indeed, platinum refuses to melt in a furnace where cast iron runs like water, and only yields to the intense heat of the oxyhydrogen blowpipe flame. (See *C. N.* pp. 5 and 6; also p. 106.) Some metals possess the very valuable peculiarity of becoming, when heated, *soft* and *pasty* before they melt, and in this condition are capable of being *welded* (from *wellen*, to join)—that is, two or more pieces, if placed together, and struck with a hammer, unite completely to form one piece. This is particularly the case with iron and platinum, and is one of the reasons why iron is more extensively and generally used than any other metal.

Many metals, and especially when cooling from the liquid state, form well defined crystals, which are generally cubical.

11. VOLATILISATION (from *volo*, I fly), or the capability of being converted into the state of vapour by heat, is believed to be possessed by every metal; the temperature,

however, at which some of them volatilise is far too high to be ascertained by our present instruments. Some of them, however, volatilise at a comparatively low temperature, and allow themselves to be distilled with ease, as is the case with mercury, which forms a colourless and transparent vapour at 350° C. (662° F.); potassium, which forms a green vapour at a red heat; zinc and others. Arsenic is peculiar, because at ordinary pressure it does not melt before it volatilises; therefore, to melt arsenic it must be heated under pressure.

12. CONDUCTION OF HEAT AND ELECTRICITY.—Of all bodies, the metals, as a class, are the best conductors of *heat* and *electricity*; their relative powers of conduction, however, vary very considerably, but, as a rule, the best conductors of heat are also the best conductors of electricity.

Let silver, the best conductor, be represented by 100, then:—

Silver	100·00	Platinum	18·03
Copper	99·95	Cobalt	17·22
Gold	77·96	Iron	16·81
Aluminium	56·06	Nickel	13·11
Magnesium	41·17	Tin	12·36
Zinc	29·02	Lead	8·32
Cadmium	23·72	Arsenic	4·76
Palladium	18·44	Antimony	4·62

13. SMELL AND TASTE.—The metals possess little smell; but certain of them, as iron and copper, emit when rubbed an odour peculiar to themselves, and arsenic when volatilised smells like garlic.

The peculiar astringent flavour of many of the soluble compounds of the metals is known as the *metallic taste.* All metallic compounds, however, do not possess a metallic taste; some, as common salt (NaCl), although metallic compounds, have a *saline taste.*

COMPOUNDS OF THE METALS.

The metals combine—

1st. With metals to form Alloys.

2nd. With the non-metals to form Metallic Salts, as Oxides, Sulphides, Chlorides, &c.

I. ALLOYS.

(From *ad*, to, and *ligo*, I bind.)

Any compound formed by the union of two or more metals is called an *alloy* unless mercury is present, when the compound receives the name of an *amalgam* (from μαλάσσω, I soften).

Many alloys are in common use, as *brass*, a compound of copper (Cu) and zinc (Zn); *bronze*, an alloy of copper and tin (Sn). They are commonly formed by melting the metals together in the requisite proportions. Alloys possess a metallic appearance, but may entirely differ in properties from one and all of their components, *e.g.* an alloy may be more fusible than any of its components. (See *C. N.* p. 7.) Some metals refuse altogether to alloy themselves with certain others; again, there are metals which will only dissolve in or alloy themselves with others in a limited degree; while others appear to be capable of alloying themselves with particular metals in any proportion. In many instances alloys are definite chemical compounds; their combinations are, however, only slightly held together by the chemical force, and in this respect show a marked contrast to the compounds of the non-metals with the metals.

II. METALLIC SALTS.

There is a great visible distinction between *Salts* and *Alloys*. In the salts of the metals the characteristic metallic properties of the metal are generally quite lost, as in the case of common salt ($NaCl$), which contains the beautiful silver white metal called sodium (Na). Further, salts may be possessed of colour, taste, smell, or almost any conceivable property and peculiarity. They may be poisonous or non-poisonous, soluble or insoluble in water or any other liquid; also fusible or infusible, decomposed or not decomposed either by heat, cold, water, or any other means. Further, they are either crystalline or amorphous, and may under one class of circumstances be possessed of one set of properties, and under different circumstances of another. Thus showing that *the properties of any compound substance can never be predicted, and cannot be known until they have been experimentally determined.* A metal may, however, by proper means, be reduced from any of its salts, and restored to its metallic appearance and characteristic properties, *e.g.* from the common white magnesia of the shops the metal magnesium (Mg) may be obtained, which can be heated at a common candle flame, when it burns with a brilliant white light to a white ash, which white ash is similar to the original magnesia.

All the metals combine with the non-metals O, S, Cl, forming the three most important classes of metallic salts, namely, the *oxides*, the *sulphides*, and the *chlorides* of the metals. Some of the metals also form chemical compounds with all the other non-metals, not even excepting H, *e.g.* the gas H_3As. The oxides and chlorides of the metals appear quite destitute of any characteristic of the metals from which they are produced; a few sulphides as *galena* (lead sulphide) and *pyrites* (sulphide of iron and copper) possess however a marked lustre resembling the metallic, while steel and cast iron, which are compounds of iron with small quantities of carbon and silicon, are remarkable for their metallic properties.

OXIDES OF THE METALS; OXYGEN METALLIC SALTS.

Although all the metals are capable of uniting with O, and many of them with several proportions of it, forming *oxygen metallic salts*, yet the affinity of the several metals for that gas is as various as may be, *e.g.*—

1. Some metals can not only rust or tarnish, *i.e.* take O from the air, but even from water H_2O, at the ordinary temperature of the air, liberating an equivalent proportion of H. (See *C. N.* p. 53.)

Metals belonging to this class are the six following; three of which are called *metals of the alkalies;* the second three *metals of the alkaline earths.*

Potassium	K	Barium	Ba
Sodium	Na	Strontium	Sr
Lithium	Li	Calcium	Ca

These metals are very light, that is, have a low s.g.

2. Other metals, again, tarnish or oxidise in the air, though more slowly than those mentioned above, but can remove O from hot water only, *e.g.* magnesium (Mg), which requires a temperature of 55° C. (122° F.) to effect the decomposition of water, while iron (Fe) and zinc (Zn) require a red heat. (See *C. N.* p. 53.) Metals of this class decompose HCl and H_2SO_4 at the common temperature with ease, liberating H. (See *C. N.* p. 52.)

Magnesium	Mg	Manganese	Mn
Zinc	Zn	Nickel	Ni
Aluminium	Al	Cobalt	Co
Cadmium	Cd	Iron	Fe

3. Another class of metals oxidise only slowly in the air, decompose water with facility when red hot, but will *not* decompose HCl and H_2SO_4 at the common temperature, but will decompose KHO (potash) liberating H. The oxides of these metals are noted for their power of entering into combination with other salts.

Antimony	Sb	Arsenic	As	Tin	Sn

4. Then there are metals which scarcely decompose water under any circumstances, yet tarnish easily in the air, and *like all the metals mentioned above,* oxidise *at a red heat* when exposed to the air. When once united to O they retain that gas with great force, and *cannot be separated from it by heat alone.*

Copper	Cu	Lead	Pb	Bismuth	Bi

5. Finally, there are those metals which, from their very slight affinity for O, and consequently from their capability to hold their lustre in the air even when heated, have received the name of *noble* metals. The oxides of these metals *do part* with their O when sufficiently heated.

Gold	Au	Platinum	Pt
Silver	Ag	Mercury	Hg

Mercury (Hg) however oxidises in the air at a low heat, but parts with its O again at a higher temperature. (See *C. N.* p. 50.) These metals have a high s. g.

The oxides of the first four groups of metals part with their O when once united to it, with as much difficulty as the ease with which they unite with it, while the noble metals part with their O as easily as it was difficult to unite them with it; yet all metallic oxides, by the use of the proper means, may have their O removed, and the metal returned or reduced to its metallic condition again.

Oxides are mostly insoluble in water, and differ so widely from one another, that little further can be said about them with profit to the student until he has separately studied them. They occur abundantly as native minerals (*ores*), frequently crystalline, from which some of the metals, as Fe, are obtained for general use.

SULPHIDES OR SULPHUR METALLIC SALTS.

All the metals are capable of forming *sulphides*. Some as copper (Cu) burn in S vapour, forming a sulphide, and most of the metals when heated in S vapour unite direct with it. Many sulphides of the metals can be formed by bringing them in contact with H_2S (dihydric sulphide or sulphuretted hydrogen), or by passing a stream of this gas through a solution of one of their salts. The formation of metallic sulphides by H_2S is a *reaction* so important that the metals may be, and for analytical purposes frequently are, classed according to their behaviour with it.

Class I.—It is found that solutions of some of the metals are not precipitated under any circumstances by H_2S, and this is because their sulphides are soluble in water.

K Ba
Na Sr Mg
Li Ca

Class II.—Again, the sulphides of some metals form double and soluble compounds with the alkaline salts potash (KHO), soda (NaHO), and solution of ammonia (H_5NO), and these sulphides are commonly said to be *soluble in hydro-ammonium sulphide* ($H(H_4N)S$) or *not precipitated in an alkaline solution.*

Sb Sn As
Au Pt

These metals can, however, be completely precipitated from their solutions by H_2S in the presence of HCl, that is, by adding HCl to the solution of the metal through which H_2S has been added, because these sulphides do not form double or soluble compounds with HCl.

Class III.—Some metals, again, cannot be precipitated in the presence of HCl, or similar H salt, because such salts dissolve their sulphides completely. These metals are precipitated completely from their solutions by H_2S, however, in the presence of an alkaline salt, with which these sulphides do *not* form double or soluble compounds. The alkaline salt usually employed to precipitate them being hydro-ammonium sulphide ($H(H_4N)S$).

Fe	Zn
Co	Mn
Ni	

Class IV.—There is a group of which chromium and aluminium are the chief representatives. Solutions of these metals are precipitated by alkaline sulphides ($H(H_4N)S$), but are precipitated *as oxides*, and *not* as sulphides.

Cr	Al

Class V.—Finally, there are those metals which are precipitated from their solutions by H_2S, either in the presence of HCl or an alkaline salt.

Pb	Ag	Cu
Hg	Bi	Cd

As a method by which *groups* of metals may be separated from one another when dissolved in the same solution, the action of H_2S is of the greatest use in the laboratory. The groups of metals precipitated by it, and the circumstances under which each group is precipitated, as mentioned, should be carefully committed to memory. This reaction of H_2S is the most important test for the metals.

The precipitated sulphides by H_2S are generally black or dark brown; zinc (Zn) however is white; arsenic (As), cadmium (Cd), and tin (Sn), yellow; antimony (Sb), orange; manganese (Mn), flesh-colour. Some salts of iron, namely the sesqui-salts, or fer*ric* salts, give first a white precipitate of S, and afterwards a black one of the sulphide; and some salts of mercury give first a white, passing through yellow and orange to black. It is, therefore, *absolutely necessary* when testing with H_2S to *add an excess of the gas*, that is, until the solution smells of the gas after it has been well stirred *or shaken, before* a reliable result can be obtained.

When a metallic sulphate is heated with C the salt becomes changed to the state of *sulphide*, the C uniting with the O contained in the salt.

Metals frequently unite with S in several proportions, forming a series of sulphur salts; the sulphur salts, however, are not so numerous, and have not been so well studied as the O salts.

The sulphides resemble the oxides very closely in their composition, and occur, frequently crystalline, as natural minerals (*ores*), from which some of the metals are principally obtained for commerce, *e.g.* copper. They often possess a kind of metallic lustre, conduct electricity, can seldom be reduced by heat alone to the metallic condition, and by roasting are generally converted into oxides.

CHLORIDES OR CHLORINE METALLIC SALTS.

The action of Cl upon the metals is generally more energetic than that of either O or S. Most of the metals unite with Cl at the ordinary temperature, and some, as antimony (Sb) and arsenic (As), &c., take fire when introduced into that gas, a chloride of the metal being formed.

Many metallic chlorides can be formed by dissolving the metallic oxides in HCl or a mixture of HCl and HNO_3, or by heating an oxide and passing Cl gas over it, sometimes with the addition of C to the oxide, CO and a metallic chloride being formed.

The metallic *iodides*, *bromides*, and *fluorides* resemble the chlorides very closely, and can be as a rule formed in a similar manner to the chlorides. The iodides and bromides, however, are very seldom ever found as native minerals.

The chlorides can generally be melted by heat, many of them being volatile, and capable of being distilled. They mostly dissolve in water without decomposition; exceptions to the above statements will be mentioned hereafter. They cannot be reduced or decomposed by heating with carbon, but many are reduced when heated in a current of H gas.

Common salt (NaCl) occurs as a native mineral; but few chlorides, however, are found among the native minerals as ores of the metals.

OCCURRENCE AND DISTRIBUTION OF THE METALS.

THE metals whose compounds are most widely distributed, and which, next to silica or sand, compose the greater part of the rocks of our earth, so far as we are at present acquainted with them, are—

Aluminium	Iron
Calcium	Sodium
Magnesium	Potassium

The other metals are only found occasionally, and distributed in isolated patches through the earth; although sometimes, as in meteoric stones composed of iron, nickel, &c., they fall from the sky.

The above metals, from their great affinity for O, S, and Cl, are, with the exception of iron, never found in the earth *native*, that is in the metallic condition, but occur in chemical combination with other substances, from which they have to be separated before they appear as metals. On the contrary, the rare and *noble* metals Au, Ag, Pt, Hg, are generally found in the native state, on account of the comparatively small affinity they possess for the non-metals. Arsenic, copper, nickel, and some other of the baser metals are occasionally found native, and frequently as alloys; *e.g.* arsenic (As), nickel (Ni), and cobalt (Co), are generally found alloyed together in certain native minerals.

Metallurgy (from μέταλλον, a metal, and ἔργον, work), or the art of working metals, is especially applied to the processes by which the metals are extracted from their ores ready for commerce.

Ores.—The chemical compounds found in the earth from which the metals are obtained for commerce are called *ores*. The principal of these are the *oxides* and *sulphides* of the metals; their carbonates and sulphates, &c., when found in sufficient quantity, are also used to yield their respective metals, *e.g.* iron occurs as *oxide* in the mineral *hæmatite*; lead as *sulphide* in the mineral *galena*; sodium as *chloride* in the mineral *rock salt*. Again, the chief source of the salts of strontium is its *native carbonate* called *strontianite*; barium is found represented by its *native sulphate* as *heavy spar*. Further, a metal is frequently found in each or all of the above combinations, the one selected from which to obtain it depending then upon extraneous circumstances, as composition, position, cost, means, &c. &c.

The ores of all the rarer metals are generally found in thin layers, called *veins* or *lodes*, imbedded in commoner rock called the *matrix* (from *mater*, a mother).

Mining consists in burrowing into the earth, or digging down to the *ore* required, separating this from the matrix as far as practicable, and sending it up to the surface.

The next process, unless the ore is already sufficiently pure, is commonly to stamp it to powder. Then, because the minerals from which the most useful metals are obtained are heavier than the matrix in which they occur, the stamped ore is placed in some kind of box or other suitable arrangement, and washed in a stream of water, which carries the lighter particles of matrix away, the heavier ore remaining at the bottom, where it is collected and preserved ready for *smelting*.

SMELTING, a process of reduction, is the operation of extracting metal from its ores. It consists in removing the bodies with which the metal is in combination in the ore, and separating and obtaining the metal itself. This process is generally effected by means of C and H, assisted by heat. Lime (CaO), or some similar substance is also commonly added to form a slag and act as a flux. (See *C. N.* p. 96.)

A SLAG is a vitreous substance, a coarse kind of glass. In *smelting furnaces* the bottom of the furnace is generally closed to hold the liquid metal when formed, the fire being maintained and urged by means of large bellows or rather blowers worked by machinery.

When the coal or fuel is burned in the same furnace in which the metal is liberated, the ash from the fuel would soon accumulate and fill up the bottom of the furnace were it not removed. The ore also, when put into the furnace, generally contains variable quantities of matrix and other impurities, which likewise require to be removed from the furnace at certain intervals.

Such smelting furnaces as are here spoken of in which the ore, coal, and lime are all mixed together, are very large, and must be kept at work for years, that is until they are burnt out, without ever allowing the fire to subside, on account of the expense and loss of time in starting them, that is getting them hot enough to properly reduce the ore.

Lime and similar bodies are used because they act as fluxes, that is they chemically unite with the ash of the fuel and with the matrix, forming a compound, namely the slag, which easily melts in the furnace heat. Thus the ash &c. from being solid and almost infusible, become, heated with the lime, a liquid, and as the slag thus formed is lighter than any of the metals which are thus obtained, the slag swims on the surface of the liquid metal, and may be run, like water, from a small plug-hole out of the furnace.

When a sufficient quantity of metal has collected at the bottom of the furnace, which the workmen know by practice, it is also allowed to run from a small hole in the side of the furnace into narrow channels made in a bed of fine sand. The plug-holes are quite small, and are stopped by pushing a plug of fire-clay, fixed at the end of a long iron rod, into the hole.

If a SULPHIDE is the ore from which a metal is to be obtained, it is generally *roasted* in the open fire, or in a kind of kiln, before it goes to the furnace. By roasting, most of the S of the sulphide burns to SO_2, while the metal generally becomes converted into an oxide.

If an *oxide*, or roasted ore, is the ore to be reduced, heat and fuel (C and H) are generally used. The greater the heat, the greater appears to be the affinity of C for O (see *C. N.* p. 60); and hence the great use of coal, charcoal, and coke in all smelting works.

There is no *native chloride* that serves for the direct production of any metal in large quantities, *e.g.* silver is occasionally found as native silver chloride, called *horn silver.*

THE REVERBERATORY FURNACE (from *re*, back, and *verbero*, I beat) is a furnace used both for roasting and reducing ores which do not require a slag. In these furnaces the fire is separated from the ore by means of a partition or bridge, and the flame and heat only allowed, *by reverberation*, to come in contact with it. These furnaces are commonly furnished with doors on each side. When in use, if these doors are opened, air gets in and the furnace becomes a roasting or oxidising furnace; if they be closed, a reducing furnace. The bed of the furnace, where the ore lies, is horizontal and slightly hollow, like a tea-saucer, to hold the liquid metal. The bed is sometimes round, but generally in form of a parallelogram, and is *fed*, that is the materials are introduced, from a hole in the top. The fire is placed at one end and the chimney at the other. These furnaces cannot produce such intense heat as the smelting furnaces.

THEORY OF SALTS.

THE present theory of salts is in reality the result of many theories, held at different times and abandoned as the progress of chemistry rendered other views concerning them necessary.

The better to understand how the several parts of the present theory came to be united, and to appreciate the time and labour expended in arriving at our present position, the following short sketch of the rise and fall of the several theories which have chiefly occupied the attention of chemists since the time of Lavoisier may be acceptable.

Until the year 1786 no attempt was made to reform the nomenclature of chemistry, and indeed up to about 1660, the date of the foundation of the Royal Society of London, there appears to have been no distinction drawn even between vegetable and mineral bodies, or rather between *organic* and *inorganic* chemistry. Thus the old chemists in their classifications placed vinegar among the mineral acids because it was sour; they also placed spiritus vini (alcohol) by the side of spiritus Libavii (tin chloride); butter of antimony (antimony chloride) by the side of cow butter, because they looked alike. Names were thus given without respect to the nature of the body, but from any accidental circumstance whatever.

It was in 1786 and following years that Lavoisier and others presented to the world the foundation of our present system of chemical nomenclature, and Lavoisier was the first to attach a chemical meaning to the terms *radical and base.* He believed O to be the great active and acidifying principle in nature, and he called the body or bodies united with O the *radical* of the compound. He further divided bodies into two groups, *acids* and *bases*, and said that when these united *salts* were the result. This acid and basic constitution of salts was called the *dualistic theory*, and by the year 1795 was generally accepted by chemists. The discovery of salts which contained neither acid nor base was sufficient to eventually dispose of this theory, and it is consequently not taught in the present day.

In 1789 Higgins of Cambridge revived one of the theories of the ancient Greeks, and said that elements consisted of indestructible *particles*, and that compounds were formed by the union of these particles. Then in 1808, Dalton of Manchester, in his 'New Chemical Philosophy,' enlarged upon the views of Higgins, and propounded the present *atomic theory*, which well adapted itself to the *dualistic theory* already then accepted. Although there are objections to the atomic theory, and many celebrated men have not believed in it, yet it is the one by which chemistry is taught at the present day, and as it is required at chemical examinations the student must accept and learn it accordingly.

In 1812 Sir H. Davy, in his 'Elements of Chemical Philosophy,' suggested the present *binary theory* of salts as something quite as likely to be true as the dualistic theory of Lavoisier and others. From that time, however, until within the last few years this theory had little support on account of attention being drawn from it by a feud waged between the supporters of other theories, to be mentioned hereafter.

In 1819, Berzelius published his celebrated *electro-chemical theory*. The great array of facts which the author of this theory brought to its support and which caused it to be rapidly accepted on all hands as an article of chemical faith, had been accumulating from the labours of Davy and others since 1800, the year in which Volta first made known the instrument called the *voltaic pile.* This theory stated that as the non-metals with the exception of H were electro-negative to the metals, and all compounds then known were formed by the union of electro-positive with electro-negative bodies, therefore no chemical union or decomposition could take place except between bodies oppositely electrical. This theory agreed very well with the dualistic theory, for as there are only two electricities, positive and negative, so bodies must be built up by pairs. The discovery that electro-negative Cl could not only displace, but also be substituted for, electro-positive H gave however the death-blow to this theory.

In 1832 Berzelius published his *radical theory.* Gay-Lussac in 1815 had discovered the first compound radical, namely, cyanogen (CN), and in 1832 Wöhler and Liebig discovered the compound radical benzoyl ($C_{14}H_5O_2$); and after this Berzelius propounded, in similar words to those which Lavoisier had used in 1793, his celebrated radical theory, and he had the great satisfaction of finding all the chemical world, including Liebig and even Dumas himself, faithful believers in the theories he had promulgated. Dumas at this time (1834 to 1837) was the leading French chemist, and became Berzelius' great opponent in the chemical strife which shortly after this commenced.

In 1834 and following years the discovery was gradually made by several French chemists, Dumas, Gay-Lussac, Laurent, &c., that the electro-positive H could be removed from a compound by the electro-negative Cl, and further that the electro-negative Cl could enter the compound operated on, and take the place of the displaced electro-positive H. This discovery of the *law of substitution* led to one of the hardest fought chemical battles on record.

Of course Berzelius, at this time the greatest chemist of the age, could not consent at once to lay aside his *electro-chemical theory* which had been accepted by the scientific world upon his authority, backed as it was by a most numerous and powerful array of experimental facts in support. He laughed at the substitutionists, but at the same time altered his radical theory in order to be in a better position to combat this *theory of substitution*, which threatened, backed by Dumas, Laurent, and Gerhardt, to disprove his most cherished ideas, and teach the world a new and opposing doctrine.

The progress of the substitutionists was delayed by an interesting little quarrel between Dumas and Laurent; they eventually, however, joined in forwarding the new ideas.

In 1839 Dumas propounded the *type theory* to assist the substitution theory, which had been chiefly the work of Laurent and others.

So long as the substitutionists were divided amongst themselves Berzelius cared little for them, but about 1839, when they joined in earnest, Berzelius took up the cudgels to defend his own theories, apparently with the firm determination rather to die than ever give up the position he had earned of having given a chemical faith to the whole world. Thus commenced a battle royal between the substitution and type theories versus the radical and electro-chemical theories.

Some of the expressions used in the above description may seem unadapted to the subject, but really they scarcely convey an idea of the case, for the disputants got so far, that they more than hinted to their opponents that they thought them blind, lunatics, &c.

Unfortunately for Berzelius, Liebig, who at this time was rapidly rising to that height of greatness which has since made his name as familiar as a household word, had not the same reason for opposing the theory of substitution and holding to the electro-chemical theory which Berzelius had. Accordingly, in 1839, as facts began to multiply in favour of the law of substitution against

Berzelius, Liebig deserted and joined the substitutionists. This was a great blow to the Berzelius party, and made their chief more determined and ferocious than ever.

The result of this struggle, which lasted for years and convulsed the chemical world with strife and party feeling, was to retard the progress of free thought; to confirm the theory of chemical substitutions, which nothing but bitter party feelings could have prevented Berzelius from allowing; to cause a fusion of the two theories of types and radicals, which were never really opposed; and to show the shortcomings of the electro-chemical theory, which in its entirety is not now held by any chemist of note.

Finally, there is the nucleus theory of Laurent and the unitary system of Gerhardt, which require notice. I 1848 and 1856 Gerhardt published his *unitary* system *molecules.* His system is to look upon the atoms c posing any compound as a whole, united to form complete and single *molecule.* This view of the constitution of salts will be referred to again.

The nucleus theory, invented by Laurent in 1836, only remains in the chemistry of Gmelin, the English translation of which by the Cavendish Society comprises eighteen volumes.

The present view of the constitution of salts recognised to be taught is the one called the *Binary Theory of Salts* (from *Bini,* by twos), first suggested by Sir H. Davy in 1812. It is dependent upon the theory of radicals, and professes to look upon every compound as the result of two radicals or sets of radicals. As types (see *C. N.* p. 47) are used at present, especially in organic chemistry, to show the supposed manner in which complex bodies are composed or put together, it follows that our views as at present taught are a mixture of the *radical theory,* extremely modified from the original as put forth by Lavoisier, and as taught and *altered by* Berzelius; and the *type theory* of Dumas and

Laurent; these, improved by a fusion with Gerhardt's *unitary system of molecules*, make up a union of theories —our present view of the constitution of salts.

It thus appears that beyond the one isolated fact that all bodies unite by weight, and consequently by number, and gaseous bodies in addition by measure or volume, nothing, absolutely nothing, is known for certain of the composition and constitution of chemical compounds. The student must therefore accept the facts laid down in works on chemistry explained by the prevailing theory of the day, and be satisfied that although he does not know the whole truth, yet that he knows what is for the want of something better accepted in the present state of chemical knowledge.

The binary theory of salts therefore is only one of convenience, and is accepted and taught, not as the truth, but because it is as good a theory for the purposes of teaching and learning as any other.

RADICALS.

The term radical, like the theory bearing its name, has undergone changes in its application since it was first used, and is, or may be, applied to nearly any substance that is susceptible of chemical action or decomposition. Radicals are divided into several classes, as follows:—

1st. Simple Radicals.—These are the elements, *e.g.* Cl, O, S, H, Ag, K, Ca, &c.

2nd. Compound Radicals.—This is a most important group of bodies, especially in the study of organic chemistry; so much so, that organic chemistry has been defined to be the chemistry of compound radicals. They are extremely numerous, and include all such bodies as cyanogen ($Cy = CN$), which, although they are compounds, the other 12 non-metals are the *non-basylous* radicals—namely, O, N, C, S, P, Se, Cl, Br, I, F, Si, Bo. Compound radicals are grouped in a similar manner; their relative position, however, must be learnt as they occur.

CLASSES OF SALTS, AND QUANTIVALENCE OF RADICALS.

The number of salts at present known is very large, and their composition very various. It is found, however, that many salts possess some peculiarities in common, which enable the chemist, in a limited degree, to arrange them in certain classes and groups, thus assisting the learner to obtain a larger and a better idea of them than he could do without some such arrangement. Thus, for example, certain salts may be arranged together, because they all contain one atom of univalent matter, capable of substitution. Yet each salt will be quite distinct from the others, as, perhaps, in colour, taste, solubility in water, in alcohol, or other particulars.

According to the binary theory, a salt is a body formed by the union of two radicals or by pairs of radicals. (See p. 22.)

Thus all compounds are looked upon as having a similar constitution to the salts of the halogens, *e.g.* common salt (NaCl), hydric chloride (HCl), which are each composed of two radicals, the first of Na and Cl, and the second of the two radicals H and Cl.

Now, if NaCl is a true salt, and composed of two radicals, according to the requirements of the binary theory, then oxides and sulphides are looked upon as salts similarly composed; *e.g.*—

$Na \mid Cl$	$Na_2 \mid O$	$Na_2 \mid S$
$Na \mid Br$	$K_2 \mid O$	$K_2 \mid S$
$Na \mid I$	$Li_2 \mid O$	$Li_2 \mid S$
$Ag \mid Cl$	$Ag_2 \mid O$	$Ag_2 \mid S$
$H \mid Cl$	$H_2 \mid O$	$H_2 \mid S$

Further, the composition of the *sulphates, nitrates, chlorates*, &c., is represented as similar to that of NaCl; *e.g.*

Na_2	SO_4	Na	NO_3	Na	ClO_3
HNa	SO_4	K	NO_3	K	ClO_3
Ca	SO_4	Ca	$2NO_3$	Ca	$2ClO_3$
Cu	SO_4	Cu	$2NO_3$	H	ClO_3
Fe	SO_4	Ag	NO_3		
H_2	SO_4	H	NO_3		

The student must be careful not to fall into the error that the salts above referred to are really constituted as represented. They are more probably constituted with all their atoms arranged together in a *single molecule*, so that their respective chemical affinities or attractions are mutually satisfied, and remain so until some new external condition, as heat or the presence of a body possessing a stronger attraction for some particular member of the molecule, operates upon it, causing a chemical change, either of substitution or decomposition; *e.g.* thus H_2SO_4 would remain H_2SO_4 until the end of time if left to itself; but the moment Zn is brought in contact with it, the H_2 is separated from the molecule H_2SO_4, the Zn taking its place, and forming $ZnSO_4$.

The quantivalence of the radicals is frequently shown by means of small *dashes* placed on the top right-hand side of their symbol, *e.g.* K', Ca'', P''', C''''. (See *C. N.* p. 39.) It is usual to place the basylous radical before the non-basylous, so that common salt is written NaCl, and not ClNa. The basylous radical, among the simple radicals, is the more electro-positive of the two. (See *C. N.* p. 50.)

Just as the elements are either uni, di, tri, or quadrivalent, so are compound radicals. And in accordance with their quantivalence (see *C. N.* p. 38), so can they give rise to certain classes of salts.

Thus the following basylous radicals, K', Na', Li', Ag', H', and the following non-basylous radicals, Cy', $(NO_3)'$, Cl', Br', I', are all univalent. Their salts will only possess a single proportion of univalent matter capable of substitution, and, consequently, will be all of one class, viewed in this light—

$NaCl$	$NaCy$	NaN_3O
KCl	KCy	KNO_3
$AgCl$	$AgCy$	$AgNO_3$
HCl	HCy	HNO_3

And we see why these radicals only form one chloride, nitrate, or cyanide, because they are univalent and, as appears in the above table, have their chemical capacities or quantivalence saturated.

The basylous radicals Ca'', Mg'', Zn'', Fe'', Pb'', and the non-basylous radicals O , S'', $(SO_4)''$, are all divalent. The salts of such non-basylous radicals with basylous radicals may be of three distinct classes—namely, they may possess one atom of divalent basylous radical, or two atoms of a univalent basylous radical, or one atom each of two univalent basylous radicals:—

$Ca''(SO_4)''$
$Fe''S$
$Zn''O$

K'_2SO_4
Na'_2S
H'_2O

$HKSO_4$
$HNaS$
HKO

The salts of the univalent non-basylous radicals with divalent basylous radicals will have the following composition:—

$Ca''Cl'_2$
$Zn''Cl'_2$

$Ca''2(NO_3)'$
$Pb''2(NO_3)'$

$Cu''Cy'_2$
$Pb''Cy'_2$

The trivalent radicals, as P''', As''', Sb''', can give rise to a still larger number of distinct classes of salts, because the three atoms of univalent matter with which they are capable of uniting may be of one kind, of two, or of three kinds. Further, it may be replaced by either divalent or trivalent matter, thus giving rise to an almost endless variety of salts:—

$H'_3(PO_4)'''$
$H'_2Na'PO$
$H'Na'_2PO_4$
$H'(H_4N)'Na'PO_4$

$H'Ca''(PO_4)'''$
$(H_4N)'Mg''PO_4$
Na'_3PO_4
$Bi'''PO_4$

$Ca_3''(PO_4)'''_2$

Hydrates.—Water, that is di-hydric oxide (H_2O), is by far the most important chemical salt, and takes part in more compositions and decompositions than any other compound. It is necessary to draw a distinction, however, between di-hydric oxide (H_2O) as a *chemical agent*, and *aqua*, simple water, as used for washing, dissolving, cleaning, &c., when it is used only as a *physical agent*.

It is found that many bodies, as di-potassium oxide (K_2O), and calcium oxide (CaO), &c., when they come in contact with H_2O, either in the liquid or gaseous state, instantly unite with it, forming definite chemical compounds, which are dry, white, and solid. It is evident, from their being dry, white, and solid, and not decomposed at 100° C., that the H_2O they contain does not exist in them in the form of aqua, yet H_2O may have been added to them direct in the form of water. These and similar bodies have received the name of *hydrate* (from ὕδωρ, water), because formerly they were supposed to be compounds of *water* with the *oxide, e.g.*—

$$K_2O + H_2O = 2KHO; \text{ and } CaO + H_2O = CaH_2O_2$$

There is, however, no direct evidence to show that either of these compounds contains di-hydric oxide (H_2O) although they contain H_2 and O.

Anhydrate (from ἀ, not, and ὕδωρ) is a term applied to bodies which are sometimes united with H_2O, and sometimes not, to distinguish them when *not* in union with that molecule, *e. g.* CaO is anhydrous calcium oxide, while CaH_2O_2 (CaO,H_2O) is hydrated calcium oxide. $CuSO_4 + 5H_2O$ common crystals of cupric sulphate; $CuSO_4$ is the anhydrous salt.

Water of Crystallisation.—Most salts are capable of crystallisation (see *C. N.* pp. 29, 31), and many of them when crystallising out of a solution in water, or from a solution containing water, take up a variable proportion of it; not exactly in chemical, but apparently in physical combination, to enable them to form the peculiar geometrical figure in which the salt crystallises. This water is called *water of crystallisation*; it can be separated again from the crystal by heat, usually without decomposing the salt itself, but the form of the crystal is generally destroyed by the process, *e. g.* the sulphates of iron, zinc, magnesium, nickel, and cobalt, all crystallise with the same quantity of water, namely, 7 molecules ($7H_2O$), *e. g.*—

$$FeSO_4+7H_2O \qquad MgSO_4+7H_2O$$
$$ZnSO_4+7H_2O \qquad NiSO_4+7H_2O$$

This $7H_2O$ can be removed by heat, the anhydrous salt remaining. The seventh molecule of H_2O, however, is far more difficult to remove by heat than the first six. Further, some other salt, as K_2SO_4, may be substituted for this seventh molecule, thus, $MgSO_4+K_2SO_4.6H_2O$. A molecule of water acting in this manner is called *basic*, *basylous*, or *salifiable* water, to distinguish it from the ordinary water of crystallisation.

Again, the class of salts called alums generally crystallise with 24 molecules of H_2O as water of crystallisation. The whole of this may be removed by heat, in which case burnt or anhydrous alum remains.

Double Salts.—It is found that many salts are capable of uniting with certain others to form double salts, much as if they were compound radicals. That is to say, they will, when their solutions are mixed, crystallise out of such solutions together, in a definite manner. Such salts are the alums. The composition of common alum is aluminium sulphate (Al_23SO_4) and potassium sulphate (K_2SO_4). When solutions of these salts are mixed, alum crystallises out, which can be represented by $K'_2Al'''_24(SO_4)''+24H_2O$, or $K_2SO_4+Al_23SO_4+24H_2O$. In this case we *know that* alum contains the elements K, Al, S, O, H, and further we know that the relative proportions in which these elements occur is K, Al, S_2, O_{20}, H_{12}, and from these data we may construct a number of formulæ to represent what we think is the proper formula for alum. But we do not know in what way these radicals are united. For instance, there is no direct proof that alum contains a distinct body K_2SO_4, although K_2SO_4 was used to produce it. The elements of alum are supposed to be united together to form one chemical molecule, but the manner and order in which they are arranged to form the molecule is not known to chemists. (See *Hydrates.*)

BASIC SALTS.—There exists a number of salts often impossible to crystallise, which are commonly insoluble in water, and of more or less doubtful composition, to which the term *basic* has been applied, because they contain a larger amount of *basylous* matter in proportion to the non-basylous than do the regular and normal salts. Many such compounds are to be met with in chemistry, and cannot be arranged under any system at present in use. Some of these are probably mixtures in various proportions of several salts, as their composition frequently alters according to the circumstances under which they are made.

The quantivalence of many of the metals appears to be constant, and creates little doubt as to the correctness of our views, *e.g.* K, Na, Li, Ag, are univalent (see *C. N.* p. 39). The quantivalence of others, however, is very doubtful, owing to the curious combinations into which some of them enter, often rendering it difficult to say with certainty whether a metal is divalent or sexivalent, or whether it may not sometimes act as a divalent element, and sometimes as a sexivalent one. Thus, supposing a metal (Fe) is said to be sexivalent, it does not follow that it will at all times show or enter into combinations to the stated full extent of its quantivalence. Part of its power may in certain combinations be as it were *latent*, *e. g.* iron is sometimes *divalent* as in $Fe''O''$, and $Fe''Cl'_2$; sometimes *trivalent* as in $Fe'''_2O''_3$, and $Fe'''_2Cl'_6$; sometimes *sexivalent* as in $Fe^{vi}O''_3$, and similarly with the metal chromium (Cr).

Again, both these metals form compounds having the following composition, $Fe_3O''_4$ and $Cr_3O''_4$; the quantivalence here is in the proportion $(8 \div 3 = 2\frac{2}{3})$ of $Fe^{2\frac{2}{3}}$, $Cr^{2\frac{2}{3}}$, thus, $Fe^{2\frac{2}{3}}{}_2O''_4$. As such a quantivalence with fractional numbers is not allowed in chemistry, it is most probable that Fe_3O_4 is a double salt composed of $Fe''O''.Fe_2'''O''_3$, and similarly $Cr''O''.Cr'''_2O''_3$. This is further shown as probable because the common ore of chromium is $Fe\,Cr_2O_4$, which may be written $Fe''O''.Cr'''_2O''_3$. From this it appears that iron and chromium are equivalent to one another, but whether they are uni, di, tri, quadri, or sexivalent is a matter of opinion.

Again, antimony (Sb) and arsenic (As) are sometimes trivalent, as in $Sb'''Cl'_3$ and $As'''Cl'_3$, and sometimes quinquivalent, as in $Sb^{v}_2O''_5$ and $As^{v}_2O''_5$.

CHEMICAL FORMULÆ.

Hieroglyphic Formulæ (from ἱερός, sacred, and γλύφω, I carve) were used by the ancients to represent the seven metals known to them. Thus:—

Gold	was	the Sun,	represented	by	☉
Silver	"	the Moon	"	"	☽
Mercury	"	Mercury	"	"	☿
Copper	"	Venus	"	"	♀
Iron	"	Mars	"	"	♂
Tin	"	Jupiter	"	"	♃
Lead	"	Saturn	"	"	♄

These metals were supposed to possess certain mystic *relationship* with the planets. Thus copper was believed to be composed of gold ⊙ and a body they called acrimony ✠, and so its symbol became Venus ♀.

At the end of the last century Lavoisier and others introduced a new nomenclature and new symbols. Many attempts have been made to improve the present symbols and manner of writing them so as to express more completely the composition of bodies, and the arrangement of the atoms in them. Thus there are:—

Glyphic, or Glyptic (from γλύφω), Graphic (from γράφω, I write), and Pictorial, or Descriptive Formulæ.

Dalton, in 1808, used pictorial symbols in his formulæ to illustrate what he imagined his atoms to be, and how he thought them united with one another.

Glyphic or glyptic formulæ are formulæ illustrated by models to represent the atoms, in the manner they are supposed to exist in the compound. They are chiefly used in organic chemistry. In *graphic* formulæ the symbols are written within a circle, and from the circle radiate certain straight lines, their number and the way they are united showing the chemical *valence* or value of each atom. Although pictorial and other formulæ are sometimes useful to illustrate a particular view of a decomposition, or special and novel representation of ideas, yet they are not found necessary in the generality of cases; the common equation form of writing the ordinary symbols being usually sufficient to convey the necessary knowledge for beginners. Many kinds of symbols have been invented, but none have superseded those commonly in use.

Inasmuch as the constitution of salts is not known, and the existence of either atoms or molecules has never been proved, the *proper* method of writing any formulæ is purely a matter of custom and opinion. The opinions of different persons vary as to the method of writing some of the commonest formulæ. It is, consequently, found advisable to write all chemical formulæ, as far as possible, without regard to any of the various opinions respecting the internal constitution of the salts represented, and these formulæ are called *synoptic.*

SYNOPTIC FORMULÆ (from σύν, with, and ὄψις, sight) are the kind of formulæ most in favour at the present day. They profess to show the several elements and the number of atoms of each which a molecule of a body contains, without entering into the manner in which the atoms are arranged in the said molecule, *e.g.* H_2SO_4 is a synoptic formula for dihydric sulphate. On the assumption, however, that dihydric sulphate contains water the formula should be $H_2O.SO_3$, and if it be supposed to contain sulphuric dioxide (SO_2) and hydric peroxide (H_2O_2) its formula should be $2HO.SO_2$, and if it contain dihydric sulphide its formula must be $H_2S.O_4$. Thus, by adopting the synoptic formula H_2SO_4, the composition of a molecule of the body is given without any reference to the opinions of this or that chemist on the subject.

TESTING.

TESTING is the art of discovering the composition of any given substance, and a *test* (from *testis*, a witness) is any substance, act, or operation that can assist us in our endeavours.

If the substance to be tested be a solid, the first thing to be done is to observe whether it be of a metallic appearance or not; whether it be crystalline or amorphous (see *C. N.* p. 29); also its colour, hardness, and general appearance: *e.g.* if a body be crystalline, it is known that it cannot be any of those bodies which never crystallise, as *flour*; if it be very hard we should not expect it to be *butter*, &c.

TESTING IN THE DRY WAY.—The next process is to observe whether the body is altered by heat or not, and if so, what alteration does it undergo; namely, is it volatilised entirely or in part; does it char; does it melt; does it change colour? &c.

Bodies are usually heated upon a piece of platinum foil; in a small glass tube; or before the blowpipe flame, either on a piece of charcoal or in a little ring of platinum wire, the ring being filled with a melted globule of some flux. (See *C. N.* p. 96.)

The chemist must experiment until he finds out, or in the case of beginners they must learn in books, what substances act in any or each particular manner; and thus armed with knowledge, and observing what substances tell them by the changes they undergo under different circumstances, they can discover what these substances are.

To discover what radicals are contained in any substance is *qualitative* analysis. To weigh, measure, or otherwise determine the amount of each ingredient, after we have determined what substances are present, is called *quantitative* analysis.

TESTING IN THE WET WAY BY SOLUTIONS.—It is usual if, after observing the peculiarities of a body as above described, we do not know what it is, to endeavour to dissolve it. The greater part of all testing and analysis consists in dissolving and then precipitating from solution, in some altered form, the substances under examination. By the term *precipitation* is meant the sudden separation of a solid body from a solution (from *præceps*, headlong).

To learn and commit to memory what substances are soluble in certain liquids, how they may be precipitated, the colour of the precipitate, and in what it is soluble or insoluble, is a most important branch of chemical study.

Simple Solution.—The liquid first employed to effect the solution of a solid is WATER, because water is capable of dissolving a larger number of substances than any other known liquid; further, it appears to dissolve them, with few exceptions, without changing the chemical properties of the bodies dissolved (see *C. N.* p. 30), on which account a solution in water is called simple solution. Alcohol is frequently used to dissolve substances, especially resinous matters. Solutions of bodies in alcohol are commonly called *tinctures.*

Solubility of Salts in Water.—The salts of the *univalent* non-basylous radicals are, as a rule, soluble in water. Nitrates, nitrites, iodides, bromides, chlorides, fluorides, chlorites, chlorates, bromates, &c., are all, with few exceptions, soluble in water. On the contrary, the *multivalent* non-basylous radicals are to a very great extent insoluble. Thus the oxides, sulphides, carbonates, phosphates, sulphates, borates, arseniates, and silicates, are commonly insoluble. There is, however, no such thing as absolute insolubility in water. For it must be borne in mind that just as there are three chlorides unlike the rest of the chlorides remarkable for their comparative insolubility in water, so, while sodium sulphate is very soluble in water, barium sulphate is one of the most insoluble salts known. Nevertheless barium sulphate is to a certain extent soluble, as the *stalactites* (from σταλάσσω, I drop) found in Derbyshire, and sold under the name of *petrified wood* (from πέτρος, a stone), prove. For these stalactites are composed of this insoluble sulphate, which during ages has been carried by water through the rocks, and as it dripped from the roofs of the caverns, has gradually deposited this, its insoluble burden, until by little and little it has formed *these wonderful* examples of its mighty power.

Chemical Solution.—If a body under examination is found to be insoluble in water, something is added to effect a change in it, either by combining the substance with some new radical, or by removing some radical from it, so as to form a substance that is soluble. (See *C. N.* pp. 30, 69.)

Again, some bodies, although insoluble in pure water, are soluble, apparently without suffering change, in water having already some salt dissolved in it, *e.g.* hydrated calcium oxide (CaH_2O_2) is more soluble in a solution of sugar than in pure water; certain salts of magnesium are soluble in a solution of ammonium chloride although insoluble in water. Further, the presence of some particular salt in solution is able to prevent the solution of salts otherwise soluble, *e.g.* in the estimation of P, the presence of a little ammonia in the wash water prevents the solution of the ammonio-magnesium phosphate.

The salts first used to effect the chemical solution of a compound are the three following; solutions of which in water are always kept ready in every laboratory on account of the frequency of their use, namely :—

1. Hydric Chloride	HCl
2. Hydric Nitrate	HNO_3
3. Dihydric Sulphate . . .	H_2SO_4

Each of these salts of hydrogen possesses in a most marked degree the power of exchanging its H for some other basylous radical, *e.g.*—

Hydric chloride generally forms a chloride with the basylous radical of the body dissolved :—

$$Ca''(CO_3)'' + 2HCl = Ca''Cl'_2 + H'_2O + CO_2$$

Hydric nitrate usually forms a nitrate; it however may only form an oxide :—

$$CaCO_3 + 2HNO_3 = Ca2NO_3 + H_2O + CO_2$$

Dihydric sulphate similarly forms a sulphate :—

$$CaCO_3 + H_2SO_4 = CaSO_4 + H_2O + CO_2$$

Other salts, either in solution or alone, are often used to effect the solution of bodies when the hydrogen salts mentioned above fail to do so. Thus potassium nitrate (KNO_3) will attack many bodies that HNO_3 will not touch. Hydro-potassium sulphate ($KHSO_4$) will render soluble certain bodies upon which H_2SO_4 has little or no action. Although, therefore, solutions of H salts or the salts themselves are commonly used to effect the solution of other bodies because they are plentiful and cheap, and when they act, act more rapidly and easily than other salts, yet the student must not on this account consider them as a separate class apart from other salts.

Just as in a chemical point of view dihydric oxide (H_2O) and dihydric sulphate (H_2SO_4) are similar in composition, because they are each capable of exchanging H_2 for other basylous radicals; so caustic potash or hydro-potassium oxide (KHO), and hydro-potassium sulphate ($KHSO_4$), are similar in composition because each has only one H to exchange. The following table will show the analogy between H_2SO_4 and H_2O :—

$$\begin{array}{llll} H_2(SO_4) & HK(SO_4) & K_2(SO_4) & Ca(SO_4) \\ H_2(O) & HK(O) & K_2(O) & Ca(O) \end{array}$$

The principal tests used to precipitate the metals from their solutions are the alkalies potash (KHO), soda ($NaHO$), and ammonia (H_3N); their carbonates, and dihydric sulphide (H_2S).

CLASSIFICATION OF METALS.

THE classification and grouping of the metals are to a great extent matters of opinion according to the point *from* which they are viewed. There are, however, certain marked distinctions which compel some of them to be classed together, leaving others doubtful; thus—

1. There are a number of metals all possessing a *s.g.* below 5. These may be called the *light metals* :—

Potassium
Sodium
Lithium
Rubidium
Cæsium
Barium
Strontium
Calcium
Magnesium
Aluminium, &c.

They will be subdivided into—

a. Metals of the alkalies
b. Metals of the alkaline earths
c. Metals of the earths

2. The remaining metals will receive the name *heavy* to distinguish them from the foregoing: their *s.g.* is over 5 :—

Iron
Nickel
Cobalt
Manganese
Zinc
Cadmium
Chromium
Copper
Lead
Bismuth
Antimony
Arsenic
Tin
Gold
Silver
Mercury
Platinum

These will be subdivided into—

a. Iron group
b. Lead group
c. Tin group
d. Noble metals

Many of the rarer metals are not given here, and further neither the above classification nor yet those given on pages 9 and 11 according to the action of the metals with O and S, meet all the views that are held upon the subject. They are nevertheless sufficient for beginners.

FIRST DIVISION.—LIGHT METALS.

GROUP A.—METALS OF THE ALKALIES.

THE name ALKALI is derived from the Arabic, the term *kali* probably being the name of the plant from the ashes of which potassium carbonate was first obtained; and *al* is the definite article.

Potassium	K	Rubidium	R
Sodium	Na	Cæsium	Cs
Lithium	Li	Ammonium	(H_4N)

The salts of ammonium (H_4N) resemble the salts of the metals of this group so nearly, that although itself only a compound radical, it is commonly placed amongst the metals of the alkalies.

The metals of this group all decompose water in the cold, liberating H. They are all univalent. They are all lighter than water, soft, and have a great affinity for O. Nearly all the salts of these metals are soluble in water, including their *oxides*, *carbonates*, and *sulphides*. A solution in water of either their oxides, carbonates, or sulphides turns red litmus paper blue. They form only one chloride.

POTASSIUM.

[*Symbol* K (Kalium); *Atomic Proportion*, 39; *S.g.*, 0·865; *Melting point*, 62·5° C.]

OCCURRENCE.—This metal is very widely distributed in combination with other radicals in sea and spring water, and in the sap of plants; it is also a constant component of the animal frame. Most soils contain salts of this metal, obtained from the decomposition of primary rocks containing felspar, mica, and other minerals, where it *occurs united* with silica (SiO_2).

History.—Discovered by Sir H. Davy in October 1807, by the aid of the voltaic battery of the Royal Institution in London, constructed for himself, consisting of 2,000 pairs of plates. By this battery he first showed that each of the alkalies, potash and soda, and the three alkaline earths, baryta, strontia, and lime, contained a peculiar metal.

Preparation.—A mixture of potassium carbonate (K_2CO_3) and C is made, by heating to redness in a covered vessel *crude tartar* obtained from wine casks. Tartar is potassium bitartrate ($KH.C_4H_4O_6$), it is decomposed by a red heat into a black mass, consisting of K_2CO_3 and free C. This black mixture is placed in an iron bottle provided with an iron exit tube, and heated *white* hot in a suitable furnace. The metal distils over as formed, accompanied with CO gas, and is collected in mineral naphtha, a liquid free from O:—

$$K_2CO_3 + C_2 = 3CO + K_2$$

The process of manufacture is difficult, and demands great care. If the metal is required pure, it must be several times redistilled.

Potassium can also be obtained by the action of the galvanic current on KHO, or by mixing KHO and finely divided Fe, and applying a white heat.

Properties.—A soft, silver-white metal, with a beautiful lustre. It is brittle at 0° C., and forms a green gas at a red heat. It tarnishes, that is oxidises, instantly on exposure to the air, on account of its great affinity for O, for which reason it is kept under naphtha. It burns, forming an oxide when heated in the air or O; when thrown upon water it decomposes that liquid, liberating H, and producing sufficient heat by the rapidity of its action to set the gas on fire, the K giving a violet colour to the flame.

The affinity of this metal for O is very great; it can take O from CO_2 when heated to dull redness in that gas, C being deposited. K also burns in Cl, and rapidly decomposes many chlorides; it also decomposes H_2S, taking the place of H, and is one of the most powerful chemical agents.

Potassium with Oxygen.

Potassium forms three compounds with O, only one of which is of importance, namely potash or potassium oxide (K_2O), which by combining with H_2O forms 2(KHO) potassium hydrate, common caustic potash, or potassa caustica of the Pharmacopœia.

Potash . . .	K_2O
Potassium dioxide .	K_2O_2
" tetroxide .	K_2O_4

The tetroxide is yellow, and difficult to obtain pure. When K is burned in O a more or less yellow solid is formed, according to the circumstances under which the experiment is performed. This yellow solid is a mixture of oxides, and gives off free O when dissolved in water, forming the hydrate (KHO).

POTASSIUM HYDRATE.—Caustic Potash.

[KHO = 56.]

Preparation.—The caustic potash of commerce is impure. It is usually made by dissolving one part, say pound, of commercial *pearlash* (which is an impure potassium carbonate (K_2CO_3) obtained from woodashes) in 10 lb. of water, and adding half a pound of caustic lime (CaO) slaked in 4 lb. of water.

If a stronger solution of K_2CO_3 be used the lime has no action on it.

The K_2CO_3 solution is kept boiling in an iron pot, and *the slaked lime,* calcium hydrate (CaH_2O_2) and water, gradually added. A white precipitate sinks to the bottom of the pot composed of calcium carbonate ($CaCO_3$), caustic potash (KHO) remaining in solution:

$$K_2CO_3 + CaH_2O_2 = CaCO_3 + 2KHO.$$

The pot is carefully covered to protect its contents from CO_2 contained in the air, and the contents allowed to cool and settle. The clear liquid is then drawn off from the $CaCO_3$ and rapidly boiled down in a clean iron pot, until it would set to a solid if cold. It is now usually cast into sticks for sale. The manufacture of *pure* caustic potash is a matter of very great difficulty and also expense, as silver basins must be used to boil down in instead of iron pots.

2nd method.—By burning the metal in dry O and dissolving the oxides formed in water.

PROPERTIES.—A whitish solid, hard and brittle, usually sold in sticks about the size of penholders, but sometimes in lump. Must be kept in well closed bottles because it is very deliquescent, that is, has a wonderful desire and power to absorb water and become liquid; on this account a tube filled with pieces of stick potash is often used to dry gases, which are made to pass through it. *It is very soluble in water*, and the solution is a most important test in chemical analysis, and is also used in medicine. Has a great affinity for CO_2, on which account a solution of one part of the hydrate in three parts of water is the usual method of extracting that gas when mixed with others. (See *C. N.* p. 62.) The solution has a disagreeable odour and burning *caustic* taste; it dissolves the skin and destroys most animal and vegetable matters, as cork, and when very strong, spoils clothes, and slowly dissolves glass and porcelain.

When solution of KHO is added to fat or oil the two unite and form a compound soluble in water, which is *soft soap*. The term *lye* (originally a solution of wood-ashes, and now any alkaline solution) is often given to a solution of KHO, especially by the soap manufacturers.

When melted with sand or silica, KHO unites with it to form a potassium silicate, potash glass, and if the KHO is in sufficient quantity the silicate or glass formed will be soluble in water.

If KHO is heated it melts, and if more strongly heated will volatilise, but without loss of water, showing how firmly united the several elements are, and that it is not probable that it contains water, although it has both H and O.

The chemical affinities of KHO are so powerful that it ranks as one of the most active chemical agents, and is very commonly used to effect decompositions both for the purposes of manufacture and analysis.

DIPOTASSIUM CARBONATE.—PEARLASH.

$[K_2CO_3 = 138.]$

OCCURRENCE.—The salts of potassium are very extensively used in the arts, especially for the manufacture of soft soap, glass, and gunpowder. Although the salts of K are very widely distributed through the rocks of which the surface of our earth is composed, as in mica, felspar, &c., from the gradual decay of which is doubtless the way K salts get into the soil, yet up to a recent period there was not any known method of extracting them artificially from these rocks which would repay for the expense. Further, K compounds are found in all land plants, in the soils on which they grow, and in the bodies of all animals; the quantities found, however, are small, compared to the amount of other matters mixed with them; and thus their extraction by chemical means becomes difficult and expensive. Yet there are two easily available sources by which man is enabled largely to supply himself with *these salts*—namely, from the ashes of trees, and the production of nitre in the soil of India, and other hot countries.

Salts of potassium are necessary for the growth of trees, and the roots of trees extract them from the soil on which they grow. Salts of K are not volatilised by an ordinary fire. Therefore, when wood is burned, they are found in the ashes in the form of potassium carbonate (K_2CO_3). In countries where large forests exist, the wood is frequently burnt to clear the ground; and in countries where wood is burned instead of coal for fuel, the ashes are collected, and prove the great source of potassium for the glass, soap, and other manufactures.

The supply of K salts is limited, and not equal to the demand for them; consequently, they are dear in price, compared to the sodium salts, which can be obtained in any required quantity. Of late years, several new sources of these valuable salts have been found. Thus it has been long known that vegetable food removed K salts from the ground on which it grew, and it is found that the wool of sheep contains most of these salts, eaten by the sheep in their food, and that by washing the wool carefully in warm water nearly all of them may be recovered. Wool is now specially washed to recover K salts from it.

PREPARATION.—Wood ashes are composed of a mixture of salts. Some of these are soluble in water, and some not. The most important of the soluble salts is K_2CO_3. Wood ashes are therefore mixed with water—*lixiviated* (from *lixivia*, lye made from ashes)—and the clear solution boiled down in iron pots (hence the original name *potash*, the ash from the pot); it is finally burnt white or *calcined* in a reverberatory furnace, and sold as American ash or pearlash, because the chief supply for England comes from that country.

The term *calx*, lime, and *calcine*—also *calces*, or bodies calcined—all relate to the operation of heating or roasting

$Ca2NO_3$; and that when K_2CO_3 is added to $Ca2NO_3$, that nitre and chalk are formed—

$$Ca2NO_3 + K_2CO_3 = CaCO_3 + 2KNO_3$$

Whether the nitre found in the soil in India is produced by the lightning, or from the decomposition of organic matter in the hot and dry soil, where it mixes with wood ashes thrown away by the natives from their fires, or by both these methods, is a matter of opinion.

In Sweden and other countries, nitre is artificially prepared in regular nitre plantations, where long low walls of mixture—organic matter, chalk, &c.—for producing nitre are built, screened from excessive rain, and at stated periods lixiviated in water; the insoluble part is returned, and helps to make new walls. The process goes on for years.

The artificial production of nitre in countries not able to command free commerce with India is a matter of great importance, because nitre is the only salt yet found that can be employed in making gunpowder, which is an intimate mixture of nitre, sulphur, and charcoal.

PROPERTIES.—A white solid, crystallising in long six-sided prisms or columns, which are *striated* (from *stria*, a streak or crease). It is soluble in 3½ times its weight of cold water, and tastes cooling and saline. Brown paper dipped in solution of nitre and dried, forms *touch-paper*, the continuous burning of which is due to the large quantity of O contained in the nitre. When heated, nitre melts, and, if then cast into balls, forms the *sal prunella* of the shops; if more strongly heated, it gives off part of its O, forming potassium nitrite (KNO_2); at a still higher temperature, the salt is decomposed into O, N, and potassium oxide. Like potassium chlorate, nitre is frequently used in the laboratory to effect the oxidation of substances, on account of the ease with which it eliminates O; when *thrown on* hot coals, it causes them to burn rapidly, called *deflagration* (from *de*, down, and *flagro*, I burn). Iron, zinc, and other metals also burn in it, forming oxides; some sulphides also, as that of antimony (Sb_2S_3), are converted into oxides by the action of it; the S of the sulphide, being oxidised and uniting with the K, forms K_2SO_4 (potassium sulphate).

GUNPOWDER.—For making gunpowder, the nitre must be quite pure. Ordinary nitre contains common salt and other impurities, which cause the nitre to become damp in rainy weather, while pure nitre does not become damp unless actually brought into contact with water. The purification of nitre is called *refraction* of nitre. The best gunpowder may be represented by the following formula:—($S + C_3 + 2KNO_3$), or nitre 75, sulphur 12, and carbon 13 per cent., which, after combustion, will produce potassium sulphide (K_2S), carbonic dioxide (CO_2), and free nitrogen (N); thus:

$$S + C_3 + 2KNO_3 = K_2S + 3CO_2 + N_2$$

The explosive force of the powder is due to the sudden disengagement of the gases CO_2 and N.

POTASSIUM CHLORIDE.

[$KCl = 74{\cdot}5$.]

Is found in sea-water, and sometimes in natural beds, mixed with common salt. It is valuable as a source from which other K salts may be made; otherwise it is of little use in the arts; but it is of great use to the chemist, because it forms a double salt with platinum chloride ($PtCl_4$), which is insoluble in alcohol, by means of which K may be quantitatively estimated.

POTASSIUM BROMIDE.

[$KBr = 119$.]

Made in the same way as the iodide. Resembles the chloride and iodide very closely. It is used in medicine.

POTASSIUM IODIDE.

[$KI = 166$.]

This salt is much used in medicine, and as a *test* in the laboratory.

Preparation.—Add I to a solution of KHO, when potassium iodide (KI) and iodate (KIO_3) are formed, as in the preparation of the chlorate (see p. 46)—

$$6KHO + 3I_2 = KIO_3 + 3H_2O + 5KI$$

The resulting mixture of salts is then heated, to expel the O from the KIO_3, when pure KI remains.

2. By dissolving K_2CO_3 in hydric iodide (HI)—

$$K_2CO_3 + 2HI = CO_2 + H_2O + 2KI$$

3. A mixture—of iron, iodine, and water—is made; the I and Fe unite, and form a green solution of iron iodide; this is boiled with excess of Fe, and K_2CO_3 added, when KI is produced, while a black mass of iron oxide falls to the bottom.

Properties.—A white solid, very soluble in water, crystallising in cubes, and deliquescing in the air. A solution of this salt in water can dissolve iodine, forming a dark brown solution. It is a test for some metals, as, with salts of lead, it gives a bright yellow precipitate; *with mercuric* chloride, corrosive sublimate, it gives a magnificent scarlet precipitate, which is, however, soluble in an additional quantity of iodide (called being soluble in excess of the precipitant).

Hydric nitrate (HNO_3), Cl, and a number of other bodies decompose KI, separating I. (See *C. N.* p. 52.)

POTASSIUM CYANIDE.

[KCy = 65.]

Is a salt much used by the photographers, and in the process of electro-plating with silver and gold.

PREPARATION.—By heating to redness dry potassium ferrocyanide (K_4FeCy_6) in an iron crucible. (See *C. N.* p. 99.) For commercial purposes, a mixture of ferrocyanide and carbonate (K_2CO_3) is heated until cyanide (KCy) swims as a clear liquid on the surface, when it is ladled from the crucible, and poured on metal plates to cool.

PROPERTIES.—A white, deliquescent solid, smelling like essence of bitter almonds; it is very soluble in water and alcohol. The solution easily decomposes, liberating hydric cyanide (prussic acid) HCy, and is a most powerful poison. Easily melts, and is then a powerful reducing agent, on account of its affinity for O, forming a compound, KCyO, (potassium cyanate), on which account it is frequently used as a flux, for, when heated, as, for example, with plumbic oxide (PbO), metallic lead is formed—

$$KCy + PbO = KCyO + Pb$$

POTASSIUM FERROCYANIDE.—YELLOW PRUSSIATE OF POTASH.

[$K_4FeCy_6 = 368.$]

THIS salt is formed by heating in an iron pot organic matter containing N (*horns*, hoofs, skins, &c.), K_2CO_3, and iron (Fe). (See *C. N.* p. 99.) The fused mass is lixiviated in water, and the liquid drawn off, from which solution the salt crystallises out in large yellow four-sided pyramids. The salt is much used in the arts to prepare the other salts of cyanogen, and for the production of Prussian blue; a solution of it in water is much used as a test for the salts of iron and copper.

SULPHATES OF POTASSIUM.

[K_2SO_4 and $KHSO_4$.]

THE former is seldom met with, except in double salts, while the latter, hydro-potassium sulphate, is formed during the manufacture of HNO_3 from nitre (see *C. N.* p. 67), and is also frequently used in the laboratory to effect the solution of certain bodies otherwise difficult to dissolve, for when melted it acts more powerfully than the common hydrogen salt—namely, H_2SO_4.

SULPHIDES OF POTASSIUM.

THERE are five sulphides of potassium, all soluble in water. If H_2S be passed through a solution of KHO, the gas *is rapidly absorbed*, and a compound (KHS) formed. On exposure to the air, this solution becomes yellow, owing to its oxidation, a compound KS being formed.

Liver of sulphur is an impure tersulphide ($3K_2S_3 + K_2SO_4$). When potassium sulphate is heated red-hot with coal, potassium sulphide K_2S is formed.

POTASSIUM BITARTRATE.—CREAM OF TARTAR.

[$KH.C_4H_4O_6 = 188.$]

Is a salt that deposits as a hard cake inside the casks in which new wine is stored. The salt is contained in other *fruit besides* the grape. When a hot solution of this salt is made, the salt crystallises in small crystals on the surface, originally called 'cream of tartar.' It is one of the most insoluble salts of K, and is frequently used as a test for that metal.

Potassæ tartras of the Pharmacopœia is made by boiling K_2CO_3 with the above salt, when dipotassium tartrate is formed ($K_2.C_4H_4O_6$).

Rochelle salt is formed by boiling a solution of cream of tartar with Na_2CO_3, when potassio-sodium tartrate is formed ($KNa.C_4H_4O_6$).

Tartar emetic is formed by boiling cream of tartar with Sb_2O_3, when potassio-antimonious tartrate is formed.

TESTS FOR POTASSIUM SALTS.

ALL the salts of K are soluble in water, except the bitartrate and the double platinum salt.

They generally melt before the blowpipe flame, when heated on charcoal, and are not volatile.

They give a violet colour to flame.

Dihydric tartrate, $H_2.C_4H_4O_6$ (commonly called tartaric acid, because the salt tastes sour), causes in solution of either KHO or K_2CO_3 a *white crystalline precipitate* of cream of tartar. The tartrate must be added in excess, otherwise a soluble salt only is produced. The precipitate is soluble in hot water, but insoluble in alcohol.

PLATINUM TETRACHLORIDE ($PtCl_4$), commonly called bichloride of platinum, gives in solutions of K salts, to which HCl has been added, a *yellow precipitate*, slightly soluble in water, but insoluble in alcohol.

SODIUM.

[*Symbol*, Na (Natrium); *Atomic Proportion*, 23; *S. g.* 0·972; *Melting Point*, 98° C.]

OCCURRENCE.—As widely and similarly distributed as K, salts of the two metals commonly occurring together in the same substance. This metal is found native in especially large quantities as common salt, also as carbonate, borate, nitrate, and sulphate.

HISTORY.—First obtained by Sir H. Davy, a few days after his discovery of the metal K, and by the same process.

PREPARATION.—By treating sodium carbonate (Na_2CO_3) with C in a similar way to that employed for the production of K. It is easier to make Na than K, and, as it is used for the production of aluminium (Al) and magnesium (Mg), the metal is now made in large quantities.

PROPERTIES.—A soft, bluish-white, brilliant metal, *resembling* K *both in appearance and properties*; also as a chemical agent it is nearly as powerful as K. When thrown upon water it decomposes it, liberating H, but not quite so energetically as K, for it does not set the H on fire; it is, consequently, more manageable than that metal, and, being cheaper, is more generally used in chemical operations where these metals are required. It gives a *yellow* colour to flame.

SODIUM WITH OXYGEN.

SODIUM forms two compounds with O—

Soda	Na_2O
Sodium dioxide	Na_2O_2.

Soda is the only oxide of importance, as when dissolved in *water it forms* caustic soda or sodium hydrate (NaHO).

SODIUM HYDRATE.—Caustic Soda.

[$NaHO = 40$.]

Preparation.—Can be obtained from sodium carbonate (Na_2CO_3) by the same means as caustic potash from K_2CO_3.

For commercial purposes, caustic soda is now largely prepared from the 'black ash' liquor obtained in the manufacture of sodium carbonate.

Properties.—Resembles caustic potash. It is very largely used in the manufacture of *hard* soap, for, when boiled with fat or oil, the two unite and form a soda-soap, which is soluble in pure water, but insoluble in salt and water. Soap made with cocoa-nut oil and soda is, however, soluble in salt and water, and is called marine soap, because one can wash in sea-water with it.

DISODIUM CARBONATE.—Common Washing Soda; Scotch Soda.

[$Na_2CO_3 + 10H_2O = 106 + 180$.]

History.—Before the year 1823 the great source of sodium carbonate was from the ashes of sea-weed. Just as trees, that is, *land plants*, contain a large proportion of salts of K, which is found in their ash after burning them in the form of K_2CO_3, so *sea plants* contain a large proportion of salts of Na, which is found in their ash after burning, as Na_2CO_3.

Sea-weed is collected on the coasts of Scotland and other places and burned to an ash; this ash is called *kelp*, *barilla*, or *varec*.

The kelp is lixiviated with water, and the clear solution separated from the insoluble portion of the ash and boiled down. This solution contains, amongst others, two important ingredients, namely, sodium carbonate (Na_2CO_3) and salts of iodine (see *C. N.* p. 79). The Na_2CO_3 separates first from the solution as the water boils away; the salts of iodine, being *very* soluble, remain dissolved in the last portions of the liquid.

Sea-weed at the present day is burned for the iodine and K it contains, and not, as was the case before 1823, for the sake of the Na_2CO_3.

The great source of sodium is in the enormous beds of common salt (NaCl) found in many parts of the world.

To convert NaCl into Na_2CO_3 was a chemical problem of great difficulty in the earlier days of chemistry, the solution of which would prove of immense value, because *a cheap alkali* represents cheap glass, cheap soap, cheap light, cheap chloride of lime, and, consequently, cheap comforts, health, and cleanliness for the people, also cheaper and improved dyes, colours, and manufactures of a thousand kinds.

The process by which this chemical change is effected was made known about 1776, a period so memorable for the discovery of oxygen, chlorine, and many of their compounds. It was used in France, and eventually tried in England in 1814; it was little employed, however, in this country until 1823, when the government removed *the duty from salt*, and allowed to burst into rapid growth the most remarkable, successful, and important set of chemical manufactures that the world has ever seen, affording occupation to thousands of hands and employing its millions of capital, namely, the great alkali manufacture of the North (of England). And the man who invented this process, who gave it to the world perfect—for it is used to-day with the same proportions of the same ingredients as at first—who has caused such wealth to this and to other countries, such comforts to their inhabitants, and whose invention has done more to advance our general *civilisation* than perhaps any other, never reaped the benefit of his labours, but died, from blighted hope, despondency, and want, in an hospital, a miserable suicide. His name was Leblanc.

PREPARATION.—The preparation of Na_2CO_3 by Leblanc's process consists of three distinct operations:—

1. When dihydric sulphate (H_2SO_4) is poured upon common salt (NaCl), hydro-sodium sulphate ($NaHSO_4$) and HCl are produced (see *C. N.* p. 77). But if an extra proportion of NaCl be used and heat employed, the $NaHSO_4$ is further decomposed, more HCl given off, and di-sodium sulphate (Na_2SO_4) formed:—

$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl.$$

To produce Na_2SO_4 (called *salt cake*) is the first part of the process in the manufacture of the carbonate.

The operation is performed in large shallow circular iron pots, placed on brickwork over a fire, and capable of holding several hundredweights of salt and the necessary quantity of H_2SO_4 at a time. The pot is arched over, and the HCl as it escapes is carried along a flue and made to pass through lofty chambers, commonly round, filled with moist coke, on which water is allowed slowly to trickle. Here the HCl is absorbed, and the liquid which runs from the bottom of the chambers is the *common muriatic acid* of commerce.

2. It has been stated on p. 13 that when a metallic *sulphate* is heated red-hot with C a *sulphide* is the result; thus—

$$Na_2SO_4 + 4C = Na_2S + 4CO.$$

If, instead of heating Na_2SO_4 and C alone, some chalk, calcium carbonate ($CaCO_3$), be added during the roasting operation, disodium carbonate (Na_2CO_3) is formed as follows:—

$$Na_2SO_4 + 4C + CaCO_3 = CaS + 4CO + Na_2CO_3.$$

The salt cake Na_2SO_4 is mixed with coal and chalk, and heated to redness in a reverberatory furnace. More chalk is added than is represented in the above formula, to prevent the decomposition of CaS in the after process by forming a double salt (CaO.2CaS). The melted black mass is raked from the furnace, when the roasting is completed, into iron waggons, and forms when cold the 'black ash or ball soda.'

The manufacture of black ash is the second part of the process.

3. The black ash is a mass of uncertain composition, containing a mixture of salts. The soluble part of the ash, composed chiefly of sodium carbonate (Na_2CO_3), caustic soda (Na_2O), and a little undecomposed sodium sulphide (NaS), is the only part of use to the manufacturer. The insoluble part, consisting of varying proportions of CaO, CaS and C, is, as far as is yet known, useless, and lies in immense heaps in the soda yards an incumbrance to the maker.

The separation of Na_2CO_3 (called now *soda ash*) and caustic soda from the 'black ash' is the third part of the process.

The black ash is lixiviated in water, and the clear solution boiled down, during which operation the Na_2CO_3 separates and sinks to the bottom of the pans, when it is removed by perforated ladles.

After a time all the Na_2CO_3 deposits; the remaining mother liquor (see *C. N.* p. 31), is of a red colour, called 'red liquor,' and consists of caustic soda mixed with a red salt, composed of sulphides of iron and sodium, and some other impurities. The sulphides in this liquor are oxidised by forcing air through it, and by the addition of sodium nitrate ($NaNO_3$). It is finally boiled down until it will solidify on cooling, and sold in the market as commercial caustic soda (NaHO).

PROPERTIES.—Large crystals, which effloresce in the air (see *C. N.* p. 31); if heated, they melt in their water of *crystallisation*; if the heat be continued, the whole of the water may be driven off, anhydrous Na_2CO_3 remaining, which itself melts at a red heat without decomposition. The salt has a nauseous alkaline taste, and is very soluble in water.

HYDRO-SODIUM CARBONATE.—Bicarbonate of Soda.

$[NaHCO_3 = 84.]$

If moistened crystals of washing soda are exposed to a current of CO_2, the gas is rapidly absorbed, the crystals become warm, and first a sesquicarbonate and finally $NaHCO_3$ is produced. The salt is less soluble than Na_2CO_3.

A native carbonate is found near the Black and Caspian Seas, Egypt, Thibet, and other places, called *natron* or *trona*.

DISODIUM SULPHATE.—Glauber's Salt.

$[Na_2SO_4 + 10H_2O = 142 + 180.]$

For the preparation of this salt see Sodium Carbonate.

The common name of *Glauber's Salt* has been long given to this body from the name of the chemist who is said first to have prepared it. It is an interesting salt. It is found in many mineral springs, and is used in medicine; it also forms an excellent freezing mixture when finely powdered and mixed with HCl, in which it rapidly dissolves. When gently heated over a fire it melts in its water of crystallisation, the melted salt not being hotter than one's hand if the experiment be carefully performed, showing the absorption and loss of heat by the conversion of a solid to the state of liquid (see *C. N.* pp. 5, 18); further, when a saturated hot solution (see *C. N.* p. 31) of this salt in water is allowed to cool in a glass flask carefully corked, and without being shaken or disturbed, it may be kept for days without any crystals being formed in it, but if the cork be removed and some small body be dropped into the liquid to cause a disturbance in it, the whole liquid will be suddenly formed into a mass of crystals.

SODIUM SULPHITE.

[$Na_2SO_3+7H_2O=128+126$.]

THIS salt is used in bleaching and dyeing as an *anti-chlore.* It is made by burning S in air, and passing the SO_2 formed into a solution or over crystals of Na_2CO_3. When dilute HCl or H_2SO_4 is added to this salt SO_2 is evolved. A *hydro-sodium sulphite* ($NaH.SO_3$) may also be formed.

SULPHIDES OF SODIUM

RESEMBLE the sulphides of potassium, both in number and properties. They may be prepared by analogous methods.

SODIUM CHLORIDE.—COMMON SALT.

[$NaCl=58{\cdot}5$.]

OF all the salts of sodium this is the most famous and at the same time common, as its name *common salt* shows.

1st. It is found native in many parts of the earth, often in enormous beds called rock salt, especially in Cheshire, Cardona in Spain, and Wielitzka in Galicia; it is also found near Belfast, Durham, Norwich, and other places. Of all the salt mines, that of Wielitzka is by far the most remarkable. The bed of salt in which this mine occurs is 500 miles long, 20 broad, and 1,200 ft. thick. There is a village in the mine, with a chapel, and the requisites of a small colony.

2nd. From natural salt springs. Here the salt is obtained by evaporating and boiling down the brine; *brine is a strong* solution of salt in water.

3rd. There is the sea, which contains amongst other salts a large amount of NaCl dissolved in it. The salt is commonly extracted by exposing sea-water to the sun in shallow pans and afterwards boiling down the liquid until nearly solid. The last portion of liquid, the mother liquor, contains all the most soluble salts, and especially KCl and a salt of bromine. Salt obtained from the sea is impure, and is known by the name of *bay salt.* The mother liquor resulting from the evaporation of the water of salt springs and sea-water often goes by the name of *bittern,* as the Kreutznach bittern (see *C. N.* p. 78).

Finally, salt in small quantities is found in nearly all soils, springs, and rivers. It is a constant component of our food, and is present as an essential ingredient of the animal frame.

Common salt is sometimes found clear and colourless like glass. Such specimens of salt are very valuable, as they are *diathermanous,* that is, they possess the wonderful power of allowing the rays of heat to pass through them as well as the rays of light. Glass, although transparent to the rays of light, is comparatively opaque to rays of heat, as it stops 50 % of them. Water also, although transparent to the rays of light, is quite nondiathermanous to the rays of heat, if unaccompanied with light, as from a black kettle of boiling water; and further, a body may be perfectly opaque to the rays of light and yet transparent to the rays of heat, allowing them to pass through easily (see *C. N.* p. 16).

Rock salt from the mines is impure, containing amongst other substances magnesium chloride ($MgCl_2$), a very deliquescent salt. The English salt is notably free from this body, and consequently keeps better in moist weather than foreign salt.

PROPERTIES.—White, pleasant to the taste, crystallises in cubes. Soluble in three parts of cold water, and is remarkable as being scarcely more soluble in hot water (see

C. N. p. 30). It *decrepitates* (from *de*, from, and *crepitus*, a crackling) when heated; fuses at a red heat, and volatilises at a higher temperature. It is used as an antiseptic, for seasoning food, as a glaze for stoneware and for the manufacture of British alkali (sodium carbonate).

Sodium Bromide, NaBr; Iodide, NaI; and Fluoride, NaF; are all analogous to the corresponding salts of potassium.

SODIUM NITRATE.—Cubic Nitre or Chili Saltpetre.

[$NaNO_3 = 85.$]

Found in beds several feet thick in Peru and Chili. As it is slightly deliquescent it cannot be used for making gunpowder. It is largely employed instead of, as formerly, KNO_3 for the preparation of HNO_3 (see *C. N.* p. 67). This salt is also used as an artificial manure.

SODIUM DIBORATE or BORAX.

[$Na_2B_2O_4.B_2O_3+10H_2O = 202+180.$]

This salt is extensively applied as a flux by solderers and gold and silver workers. Originally the supply came from Thibet, where it is picked up in crystals as a natural product, and sent to India tied in skin bags, hung round the necks of the herds of cattle driven by the natives over the Himalaya mountains. It is called *tincal* in India. Borax is now largely made in Tuscany, where boron trioxide (B_2O_3) issues from the ground in certain places mixed with steam. This is collected by bringing the fumes in contact with Na_2CO_3 contained in pits, called *lagoons*. It is then evaporated, crystallised, and sold.

Properties.—Forms hard, clear, crystals, soluble in twelve parts of cold water. When heated the crystals froth up, gradually lose all their water of crystallisation, and form a light feathery porous anhydrous mass, *which melts to a clear and colourless liquid at a red heat.*

This salt, like many others, possesses the power of dissolving certain other salts, and it does this in a manner and under circumstances which render it extremely valuable in the arts and as a *blowpipe flux*. This solvent action is specially shown at a red heat. When gold and other metals require to be soldered, it is absolutely necessary that the two surfaces of the metal to be joined should be quite clean. To this end the metal may be covered with borax, the solder placed in the borax and heated by the blowpipe flame; the borax melts and dissolves all *tarnish* or oxide from the gold and also from the solder, and as it is liquid and light compared to the metals, floats on their surface, thus protecting them from the action of the air.

The salt has a feebly alkaline taste and reaction.

HYDRO-DISODIUM PHOSPHATE.—COMMON PHOSPHATE OF SODA.

$$[Na'_2H'.(PO_4)''' + 12H_2O = 142 + 216.]$$

THE phosphates of sodium are a very important class of bodies to the chemist, as they illustrate in a remarkable manner the different classes of salts which a trivalent non-basylous radical is capable of forming.

PREPARATION.—By adding Na_2CO_3 to the *superphosphate of lime liquor* ($CaH_4P_2O_8$) obtained in the process of making P (see *C. N.* p. 89):—

$$2Na_2CO_3 + CaH_4P_2O_8 = CaCO_3 + H_2O + CO_2 + 2(Na_2H.PO_4).$$

PROPERTIES.—Large transparent, rhombic, efflorescent crystals, soluble in four parts of cold water. When heated they melt in their water of crystallisation, and by continued heat lose it all, $2[Na'_2H'.(PO_4)''']$ remaining; *by a red heat a further molecule of* H_2O *is expelled, leaving* $Na'_4(P_2O_7)''$; this salt is only divalent, and may be dissolved and crystallised to $Na_4P_2O_7 + 10H_2O$, called pyrophosphate of sodium (from πῦρ, fire, in allusion to the way it is made).

If trihydric phosphate ($H_3.PO_4$) be added to the common phosphate ($Na_2H.PO_4$), a salt is formed having the following composition: $2(NaH_2.PO_4 + H_2O)$, dihydro-sodium phosphate; or if HNO_3 be added to the common phosphate the same salt is produced. When this salt $2(NaH_2.PO_4 + H_2O)$ is heated it first loses the H_2O its water of crystallisation, and at a higher temperature other H_2O, leaving $Na_2H_2P_2O_7$, and at a still higher temperature another molecule of H_2O is expelled, leaving $2[Na'(PO_3)']$ a salt called sodium meta- or monophosphate. In this case we have only a univalent salt. If a hot solution of ammonium chloride be added to a hot solution of common phosphate, a salt crystallises out called microcosmic salt ($Na(H_4N)H.PO_4 + 4H_2O$), hydro-sodium-ammonium phosphate. This salt is much used in blowpipe experiments, for when heated all the volatile radicals escape, $NaPO_3$ remaining, and this metaphosphate is like borax, for when melted it dissolves many salts, and forms with certain of them characteristic coloured beads or glasses.

If caustic soda be added to the common phosphate there is produced tri-sodium phosphate, $Na'_3(PO_4)'''$.

The common phosphate gives a *yellow* precipitate with silver nitrate ($AgNO_3$), while the metaphosphate gives a *white gelatinous* precipitate with that test, and the pyrophosphate a *dense white* one.

SILICATES OF SODIUM.

THERE are many natural minerals found, containing Na and SiO_2, most of them possessing a very complicated formula. Whenever silica or sand is heated sufficiently with soda, the two melt together, *glass* being the result.

GLASS is an artificial *silicate*. The Bohemian, generally called German glass, of which chemical apparatus is made, *is hard and* difficult to melt. It is composed of sand, potash, and lime (CaO). English and French white window glass is soft, and composed of sand, soda, and lime. English flint and optical glass is composed of sand, potash, and lead (as oxide). Green bottle glass is made from impure materials, and contains iron (as oxide), which gives the colour.

Coloured glass is made by adding to the mixture of materials some special body to give the colour desired; or the colour is painted on the glass and then burnt in, melted on by the furnace. Thus 'reds' are produced by a salt of gold called 'gold purple,' or the red oxide of copper; blue by a salt of cobalt; white by tin oxide, &c.

Paste to imitate jewels is a glass composed of sand, potash, and half its weight of lead (oxide). This glass easily melts, and has a high refracting power.

In silicates the more silica and less alkali, the more insoluble is the compound and the more difficult to fuse; the more alkali and the less silica the more fusible and the more soluble is the compound, until a glass may be made that will dissolve in hot water, called *soluble glass.*

TESTS FOR SODIUM SALTS.

There is no good test for these salts. They give a yellow colour to flame, and are mostly colourless and soluble in water.

LITHIUM.

[*Symbol,* Li; *Atomic Proportion,* 7; *Density,* 0·59; *Melting Point,* 180° C.]

History.—Lithia, that is lithium oxide (Li_2O), was first obtained by Arfvedson, in the laboratory of Berzelius in 1817. The name is derived from λίθος, a stone, as it was first obtained from a hard mineral. The metal first produced in quantity by Bunsen from lithium chloride (LiCl), by the electric current. It is a rare metal, although its compounds are very widely distributed in very small quantities in rocks, spring waters, and the ashes of many plants.

The salts of lithium are chiefly obtained from the minerals called lepidolite, spodumene, and petalite.

PROPERTIES.—The metal is the lightest known, as it swims in naphtha. It is white, harder than K, but softer than lead. Can be drawn into wire, is volatile, and decomposes water like Na. It forms an oxide and other salts, which resemble those of K and Na, except that they are somewhat less soluble in water, and give a brilliant crimson colour to flame. The salts of lithium are used in medicine.

RUBIDIUM.

[*Symbol*, Rb ; *Atomic Proportion*, 85·36 ; *S.g.* 1·52 ; *Melts*, 38·5° C.]

HISTORY.—Discovered in 1860 by Bunsen and Kirchhoff, and named (from *rubidus*, dark red) on account of the colour of the lines it gives in the spectrum. It resembles K.

CÆSIUM.

[*Symbol* Cs ; *Atomic Proportion*, 133.]

HISTORY.—Discovered by the same men at the same time as rubidium. This is named (from *cæsius*, sky blue) on account of the colour of the lines it gives in the spectrum, and it was by means of the spectrum that both *these* elements were discovered. Cæsium is the most *electro-positive* of all the elements.

AMMONIUM.

[*Symbol*, H_4N; *Combining Proportion*, 18.]

The existence of the compound radical H_4N is far more assumed than proved. It may be classed amongst the compound basylous radicals.

It has never been isolated, although there is one experiment which seems almost to show its existence; namely, when an amalgam of sodium and mercury is placed in a saturated solution of ammonium chloride (H_4NCl), the mercury (Hg) swells up in a most remarkable manner to ten times its original size. If this spongy mass of mercury be analysed it is found to contain $Hg + N + H_4$, or it may be looked upon as an ammonium amalgam ($H_4N.Hg$). The spongy mass of mercury cannot be preserved, for as soon as it has ceased to grow in size it begins to decrease, giving off H_3N and H, until after a short time globules of mercury alone are left. The salts of H_4N resemble the salts of K, Na, Li just as closely as the salts of CN (*cyanogen*) resemble the salts of Cl, I, and Br, so that H_4N has as much right to be considered a basylous radical as CN has to be considered a non-basylous radical.

SOLUTION OF AMMONIA.

[$H_3N + H_2O = H_5NO = (H_4N)H.O.$]

Ammonia gas is very soluble in water. (See *C. N.* p. 65.) On heating a solution of it in water the H_3N is given off uncombined with H_2O, consequently the chemical affinity between the two is very feeble. Although $(H_4N)H.O$ may be considered as ammonium hydrate having a similar composition to KH.O, yet in this respect they differ, for no heat is able to separate H_2O from 2KHO.

Solution of ammonia is one of the commonest tests in the laboratory. *It precipitates* a great number of oxides from solution, some of which are soluble in excess of ammonia solution, *e.g.* copper.

The liquid has a powerful odour of the gas, and is called caustic ammonia from its alkaline and burning effects on the skin.

AMMONIUM CHLORIDE.—SAL AMMONIAC.

$[H_4NCl = 53{\cdot}5.]$

THIS is the most important salt of ammonium.

PREPARATION.—It is almost universally formed by bringing H_3N in contact with HCl. The two have a great affinity for one another, and combine at once, forming $H_4NCl = H_3N.HCl$, which is a white solid very soluble in water.

1.—Whenever *organic matter containing* N is distilled, H_3N is given off. Coals contain N; coals are largely distilled to produce coal gas (see *C. N.* p. 102), during the formation of which great quantities of H_3N are generated and collected or retained in the 'gas liquor.'

Organic matter is frequently spoken of as containing N or not. All organic matter contains C, and some one or more elements besides, but N is not always present. Such bodies as flesh, horns, skins, albumen, gluten, gelatine, all contain N, and as they also possess H, will when distilled give H_3N; *e.g.* the horns of the hart used to be distilled for the ammonia given off from them, hence the name 'spirit of hartshorn.' Sugar, fat, oil, alcohol, and others have no N, and consequently cannot produce H_3N.

2.—'Gas-liquor' is mixed with HCl and boiled down; or the 'liquor' is mixed with a little lime and boiled, and the H_3N as it escapes made to pass into HCl, when it is at once absorbed, forming H_4NCl.

The salt is either allowed to crystallise out of a hot *solution*—when it appears as a powder, but consists of minute crystals—or the solution is boiled dry and sublimed (see *C. N.* p. 6) in shallow pans, one inverted over the other. The chloride condenses on the inside of the upper pan, and is sent into the market in semi-round, tough, fibrous masses.

AMMONIUM SULPHATE.

[$(H_4N)_2SO_4$.]

PREPARED from gas liquor by boiling it with a little lime, and retaining the H_3N in H_2SO_4.

Used as a manure, and for rendering muslins non-inflammable by dipping them in a solution of the salt.

AMMONIUM NITRATE.

[$H_4N.NO_3$.]

THIS salt is of importance, as it is the one from which nitrous oxide (N_2O) is usually made (see *C. N.* p. 68), hydric nitrate and ammonium carbonate are mixed and the salt allowed to crystallise out. It is deliquescent and requires to be kept carefully from the air. The salt produces great cold when dissolved in water, and is often used as an artificial *freezing mixture.*

AMMONIUM CARBONATES.

THERE are several carbonates of ammonium, only one of which is important, namely the common *smelling salts* of the shops $2[2(H_4N).CO_3].CO_2$, sesquicarbonate of ammonium.

PREPARATION.—By mixing chalk ($CaCO_3$) and either the sulphate or chloride of ammonium, applying heat, and collecting the product in leaden vessels. Large quantities of free H_3N are generated at the same time:

$$6H_4NCl + 3CaCO_3 = 3CaCl_2 + 2[2(H_4N).CO_3].CO_2 + 2H_3N + H_2O$$

Properties.—A white mass, volatile, like ammonia in smell, becomes powdery on the surface after exposure to the air from formation of bicarbonate. Soluble in cold wator, but decomposed by boiling. Its solution is much used as a *test.*

HYDRO-AMMONIUM SULPHIDE.

$[(H_4N)H.S = 51.]$

This salt is commonly used as a test in the laboratory to precipitate many of the metals. (See p. 11.) It is made by passing H_2S through a solution of ammonia. At first the liquid is colourless, but on keeping it becomes yellow from absorption of O and formation of the yellow bisulphide of ammonium $(H_4N)_2S_2$ and an O salt. Other sulphides of ammonium exist, all of which are soluble in water.

TESTS FOR AMMONIUM.

The salts of ammonium are *volatile* and may be sublimed, but if the non-basylous element in the salt is *non-volatile,* as in the phosphate, the salt is decomposed by the heat, H_3N and water only going off.

Whenever lime (CaO) or, better still, calcium hydrate (CaH_2O_2), is added to any salt of ammonium, vapours of H_3N are quickly given off, which may be recognised by their smell or by test-paper. (See *C. N.* p. 65.) Caustic potash or soda act like lime, but more energetically.

Dihydric tartrate $(H_2.C_4H_4O_6)$ and platinum tetrachloride $(PtCl_4)$ give precipitates with salts of ammonium analogous to those of K.

GROUP B.—METALS OF THE ALKALINE EARTHS.

THE metals of the alkaline earths are four in number, namely:—

Barium	Ba	Calcium	Ca
Strontium	Sr	Magnesium	Mg

Before the year 1807 the existence of any of the light metals of the first division (p. 40) was not known.

Lime, which is calcium oxide (CaO), was known long before the metal Ca was discovered. It was considered to be a simple or elementary substance, because it had not, up to that time, been proved to be a compound; so, from its white, earthy appearance, and its caustic, alkaline properties, it was called an *alkaline earth*. The same term, alkaline earth, was applied to the oxides, commonly called *baryta* (BaO), *strontia* (SrO), and *magnesia* (MgO), as they were severally discovered. These alkaline earths are only slightly soluble in water, to which, however, they give an alkaline taste, and the property of turning red litmus paper blue, and yellow turmeric paper brown.

The early chemists confounded lime and magnesia together, under the term *terræ absorbendes*.

This group of metals is distinguished from group A as follows:—

1st. The metals of this group are all divalent, and heavier than water. They also decompose water far less rapidly than group A, magnesium requiring the water to be warm before it can decompose it.

2nd. Their oxides and sulphides are less soluble in water.

3rd. Their carbonates are insoluble in pure water, but are soluble in water containing CO_2 dissolved in it. (See *C. N.* p. 61.)

4th. Their sulphates—

- *a.* Barium sulphate is one of the most insoluble salt known.
- *b.* Strontium sulphate is very insoluble, but less s than barium sulphate.
- *c.* Calcium sulphate is moderately insoluble, namely about 1 grain in an ounce of water.
- *d.* Magnesium sulphate is very soluble in water.

The salts of the metal magnesium bear a great resemblance to those of zinc and cadmium, on which account these three metals are sometimes placed in a separate class, called the *magnesian metals.*

BARIUM.

[*Symbol*, Ba; *Atomic Proportion*, 137; *S.g.* 4·00.]

Occurrence.—Found in the native minerals *heavy-spar*, barium sulphate (hence the name *barium*, from *βαρύς*, heavy), and *witherite*, barium carbonate, so called from Dr. Withering, its discoverer. Both minerals are found in large quantities in the north of England, commonly near *lead ore.*

History.—This metal was first obtained by Sir H. Davy. When pure, it is said to be pale yellow. It is with difficulty obtained by passing a strong electrical current through the fused chloride; and, as it is not used for any chemical purpose, specimens of the metal are scarce. It decomposes water, and oxidises in the air very rapidly.

Barium with Oxygen.

Barium forms two compounds with O, namely—

Baryta	BaO
Barium dioxide or Peroxide of barium	BaO_2

BARIUM OXIDE.—BARYTA.

[$BaO = 153$.]

HISTORY.—Baryta was discovered by Scheele in 1774.

PREPARATION.—By heating barium nitrate ($Ba2NO_3$) to redness in a crucible, when BaO remains.

It is never made, like CaO, by heating the carbonate to redness, on account of the very high temperature required to separate CO_2 from $BaCO_3$ (barium carbonate).

PROPERTIES.—A grey mass; *absorbs* CO_2 from the air, forming barium carbonate, and also water, forming a hydrate; and, if mixed with one-eighth of its weight of water, it *slakes* very much like lime under similar circumstances.

BARIUM HYDRATE.—HYDRATE OF BARYTA.

[BaH_2O_2, or $Ba(HO)_2 = BaO.H_2O$; when crystallised, $BaH_2O_2.8H_2O$.]

Is either a white solid powder, or in crystals; it is soluble in 3 parts of boiling and 20 parts of cold water. The solution has a very remarkable affinity for, or desire to absorb, CO_2; consequently, baryta-water is frequently used as a test for CO_2, instead of lime-water; it is used in a similar manner to the latter, and forms a white precipitate of barium carbonate ($BaCO_3$).

BARIUM DIOXIDE.—PEROXIDE OF BARIUM.

[$BaO_2 = 169$.]

THIS compound is the one generally recommended to be used for the production of that curious compound, hydrogen dioxide, or peroxide of hydrogen ($H_2O_2 = 2HO$).

PREPARATION.—Either by heating BaO in a current of O gas, or in a crucible, mixed with potassium chlorate.

Properties.—Mixed with water, it forms a white solid *hydrate* ($BaO_2.6H_2O$), which is insoluble in water. When dilute H_2SO_4 is added to the hydrate, $BaSO_4$ is precipitated, and H_2O_2 is liberated, which dissolves in the water.

BARIUM SULPHATE.—Heavy-Spar.

[$BaSO_4 = 233.$]

This mineral is one of the principal sources of barium compounds. It is used under the name of permanent white, as a substitute for the paint called white lead, and sometimes to adulterate it. It is not soluble in water, nor is it acted upon by any chemical salt. When, however, it is mixed with C, and strongly heated, it obeys the law of all sulphates, and becomes converted into BaS (barium sulphide), which is decomposed easily by 2HCl, forming $BaCl_2$ and H_2S.

BARIUM CHLORIDE.

[$BaCl_2 = 208.$]

A solution of this salt in water is one of the most important laboratory tests, as it forms a white insoluble precipitate of $BaSO_4$ when added to a solution of any sulphate (say $MgSO_4$).

Preparation.—Either by decomposing *witherite* ($BaCO_3$) in HCl, or by decomposing BaS, the sulphide formed by heating $BaSO_4$ and C together, in HCl:—

$$BaCO_3 + 2HCl = BaCl_2 + CO_2 + H_2O$$
$$BaS + 2HCl = BaCl_2 + H_2S$$

Properties.—Crystallises in beautiful pearly plates ($BaCl_2.2H_2O$). Soluble in twice its weight of cold water, but is *insoluble* in a *strong* solution of HCl. When testing for a sulphate by $BaCl_2$, in presence of HCl, any white precipitate produced should be mixed with water, to see whether it is soluble in that liquid, as the precipitate may be only $BaCl_2$, and not $BaSO_4$.

BARIUM NITRATE.

[$Ba2NO_3$ or $BaN_2O_6 = 261$.]

THIS salt is sometimes used in the laboratory as a test for sulphates, when it is not desirable to use $BaCl_2$, to avoid getting Cl in the solution to be tested. It is also, as well as the chloride, used in fireworks to produce green fire. It is easily made by dissolving either *witherite* or BaS in dilute HNO_3. Requires eight times its own weight of cold water for solution. When heated it melts, at a higher temperature boils up, fumes are given off, and, finally, BaO remains.

BARIUM CARBONATE.—WITHERITE.

[$BaCO_3 = 197$.]

Is found native as a white mineral. It may be formed artificially by adding any soluble carbonate, as the carbonate of potassium, sodium, or ammonium, to any soluble salt of barium, as the nitrate or chloride. It is insoluble in water, but like calcium carbonate soluble in water containing CO_2 dissolved in it.

BARIUM SULPHIDE.

[$BaS = 169$.]

THIS compound may be formed by heating heavy-spar with C, and dissolving the mass in water. It is sometimes used as a depilatory.

OTHER salts of barium, as the Iodide, Chlorate, Sulphite, Phosphate, &c., may be formed; they are, however, of comparatively little importance.

TESTS FOR BARIUM SALTS.

All the soluble salts of barium are poisonous. The antidote is glauber salt (sodium sulphate), or Epsom salt (magnesium sulphate), because—

Any soluble salt of barium is precipitated as a *white* insoluble solid ($BaSO_4$) by a soluble sulphate, even $CaSO_4$, and of course any soluble sulphate is precipitated by a soluble salt of barium; the solution to be tested, however, *should always have a few drops of* HCl *or* HNO_3 *added to it*, as otherwise the white precipitate might be carbonate, phosphate, or oxalate of barium.

With the *soluble carbonates*, as of potassium and sodium, a soluble salt of barium gives a white precipitate of barium carbonate ($BaCO_3$).

STRONTIUM.

[*Symbol*, Sr; *Atomic Proportion*, 87·5; *S.g.* 2·54.]

Occurrence.—This element is much more scarce than barium, which, however, it very closely resembles. It is found as *strontianite*, a native carbonate, and as *celestine*, a native sulphate, so named because it is sometimes of a beautiful blue colour.

History.—First found, as carbonate, at Strontian, in Argylshire, in 1787, and mistaken for barium carbonate. In 1792, Dr. Hope pointed it out as a peculiar earth. The metal was first obtained by Davy in 1808. The method of preparation and the properties of the metal resemble those of Ba very closely.

Strontium with Oxygen.

Only one oxide of strontium has as yet been formed, namely, Strontia (SrO), which closely resembles the corresponding barium oxide.

STRONTIUM NITRATE.

[$Sr(NO_3)_2 = 211.$]

THIS salt may be formed similarly to the barium nitrate. It forms crystals, containing $Sr(NO_3)_2.5H_2O$, which are soluble in five parts of cold water.

This salt, dried and free from H_2O, is much used by firework-makers to produce a splendid red fire.

TESTS FOR STRONTIUM SALTS.

THEY give a red colour to flame, which must not be confounded with the colours produced by lithium and calcium under similar circumstances. These salts are not poisonous.

Carbonates of potassium and sodium give white insoluble precipitates in salts of strontium, which are, however, soluble in water containing CO_2.

Soluble sulphates give a white precipitate with salts of strontium, but the precipitate does not show itself immediately in dilute solutions, owing to the partial solubility of the strontium sulphate.

CALCIUM.

[*Symbol*, Ca; *Atomic Proportion*, 40; *S.g.* 1·578.]

OCCURRENCE.—Next to silica and alumina, the salts of calcium, principally the carbonate and sulphate, are the most abundant minerals. They form whole mountains and deposits as chalk, marble, mountain limestone, blue lias, &c. ($CaCO_3$), and gypsum ($CaSO_4$). The phosphate ($Ca_3P_2O_8$) forms the solid part of bones; the chief part of the shells of eggs, crabs, oysters, and other crustacea being composed of $CaCO_3$.

History.—First obtained by Sir H. Davy in 1808. It is a yellowish, light metal, malleable, and harder than lead. It soon tarnishes in the air, and burns with a brilliant white light when heated in it. The metal may be obtained by decomposing one of its salts with a powerful electric current. Black, in 1756, first showed the difference between burnt and unburnt lime. (See *C. N.* p. 61.)

CALCIUM OXIDE.—Quick-Lime.

[$CaO = 56.$]

This substance has been known from the earliest times. It is obtained by heating the native carbonate to redness.

Lime Burning. Lime Kilns.—Lime burning is carried on to a very great extent in many parts of England, and other countries where limestone is found, because *lime* (CaO) and *sand* (SiO_2) form mortar. The process of lime burning depends upon the fact that, when any form of calcium carbonate ($CaCO_3$), as marble, limestone, or chalk is heated red-hot in a current of air, CO_2 is given off, CaO, as a white or grey mass, according to the purity of the material used, remaining ($CaCO_3 = CaO.CO_2$). In the laboratory, this may be accomplished by burning pieces of marble in a crucible with a hole at the bottom; at the lime-pits, limestone is burnt in *kilns*, which may be termed *large* crucibles, with the fire lighted inside them, at the bottom; layers of stone and coal are repeatedly heaped on the top of one another, volumes of CO_2 escape, and the lime (CaO), as it is ready, raked out at the bottom, so that the process is a continuous one.

Properties.—This body possesses a great affinity for water, with which it forms a solid hydrate. On account of its power of abstracting water, it is frequently employed in the laboratory to dry gases, and deprive other bodies also of moisture. When mixed with cold water, it becomes hot, owing to the chemical action between the water and the CaO; the mass cracks, and falls to powder, the powder occupying more space than the original lime, on which account, as the lime appeared to be alive, from thus cracking and increasing in size, it has received the name of *quick-* or *live*-lime; and the process of throwing water on it is called *slaking*, the lime appearing to be thirsty, owing to the quantity of water it absorbs or drinks up; so, when its thirst is quenched, it is said to be *slaked*, and, after it has fallen to powder, it is often called *slacked* lime, that is, loose or powdered lime. When once united with water, a similar action cannot again be brought about, unless the powder be reheated. *Quick-lime* is very caustic and burning to the skin; *it cannot be melted* when pure, and, as it emits a beautiful white light when intensely heated, it is used for the *lime-light*, which consists of a small cylinder of well-burnt CaO, heated to intense whiteness by the oxy-hydrogen gas flame. It is also made into crucibles, for use with the same flame; in these platinum may be melted.

CALCIUM HYDRATE.—SLAKED LIME.

[CaH_2O_2, or $Ca(HO)_2 = CaO.H_2O$.]

FORMED by pouring water upon well-burned CaO, when the mass swells up, steams, and, if too much water has not been added, a light white powder is formed. The heat produced by the chemical union of the CaO and H_2O is so great that meat may be cooked by it, and workmen commonly warm their food by slaking a little lime, and placing the vessel containing the food over it, as if over a fire. If a larger quantity of water be added, and the mass stirred, a white sludge is formed, called *lime-wash*, or milk of lime. If a small quantity of lime is shaken with a still larger quantity of water, and allowed to settle, the clear liquid forms *lime-water*, used as a test for CO_2. (See *C. N.* p. 61.) Lime-water, if exposed to the air, soon becomes covered by a thin film of a white substance, which is $CaCO_3$, formed by CO_2 absorbed from the air by the lime-water. Lime-water contains about 1 part of CaO in 700 of water, or half a grain to an ounce. The solution has an alkaline taste and reaction.

If CaH_2O_2 be heated red-hot, H_2O is given off, CaO remaining. It is owing to the great power CaO possesses of abstracting water and CO_2 from the air that *quick-lime* cannot be long preserved, unless kept in closely covered vessels.

Mortars and Cements.—The chemistry of the difference between mortars and cements, hydraulic limes (from ὕδωρ, water, because they will set firm even in water), and others, is obscure.

If a lime is nearly pure, it is called a *fat* lime. If it contain a large quantity of impurity, as sand, clay, &c., it is called a *poor* lime.

Mortar is lime and sand mixed together, and afterwards worked to the consistency of a stiff mud with water.

Cement is lime and clay; it appears to be a necessity that the clay should contain free silica and oxide of iron. There must be at least 15 % of clay in a lime to form a cement. *Roman cement* is a poor lime, which contains the proper quantity of clay, &c. to enable it, after being burned, to set rapidly under water.

Lime is further extensively used as a manure; in the purification of coal-gas; to remove the hair from hides, in the process of tanning, on account of its caustic properties; and in the manufacture of a number of other articles.

CALCIUM CARBONATE.—CARBONATE OF LIME.

[$CaCO_3$ or $CaO.CO_2 = 100$.]

THIS mineral is found in some parts of the earth in enormous masses. Thus, as *limestone*, it forms very extensive beds of deposit in the earth, as well as great rocks out of it; and, as chalk, it composes the *downs* and *cliffs* of the south of England. When crystalline, it forms all the different kinds of *marble*, and further is the principal component of the geological formation called the *oolite* (from ὠόν, an egg, and λίθος, a stone), calcareous marl, mountain limestone, lias, shells of all kinds, and corals. It is seldom met with pure. It forms an extraordinary number of crystals, of which dog-tooth spar and the six-sided prism are perhaps the most common English forms; it is therefore dimorphous. (See *C. N.* p. 29.) As Iceland spar, it is most invaluable to the optician, owing to the peculiar way it acts on rays of light passing through it, and, as a *pearl*, is very valuable to the jeweller.

It is formed whenever a *soluble carbonate* (say Na_2CO_3) is added to a soluble salt of calcium (say $CaCl_2$).

Properties.—Calcium carbonate is decomposed by the three H salts, HCl, HNO_3, or H_2SO_4, with escape of CO_2, and formation of a corresponding compound with the non-basylous radical of the H salt (see *C. N.* p. 61):—

$$2HNO_3 + CaCO_3 = Ca2NO_3 + H_2O + CO_2$$

It is decomposed by a red heat in a current of air, with escape of CO_2. (See *Lime burning.*)

It is insoluble in pure water, but soluble in water containing CO_2 in solution; this accounts for chalk being found dissolved in nearly all natural waters. Water containing large quantities of chalk dissolved in it deposits a portion of this if allowed to trickle over a rough surface, or slowly to drop from a small hole; in the former case, the effects of the petrifying wells of Derbyshire and other places are produced; and in the latter, the stalactites and stalagmites with which many caves are adorned.

Waters containing $CaCO_3$ in solution have the whole of it precipitated by boiling them; this drives off the CO_2, by the action of which the $CaCO_3$ is kept in solution.

(See *C. N.* p. 57.) The $CaCO_3$ may also be precipitated by the addition of the proper quantity of lime-water to unite with the CO_2 contained in the water itself, when the whole lime is precipitated as insoluble $CaCO_3$.

CALCIUM SULPHATE.—GYPSUM.

[$CaSO_4 = 136.$]

THIS mineral is found in very large quantities in Derbyshire, Paris, and some other localities. It is a common constituent of spring water (see *C. N.* p 57), causing *permanent* hardness.

Gypsum, as dug out of the earth, usually contains two molecules of water of crystallisation, $CaSO_4.2H_2O$. When heated to about 130° C. (266° F.), it loses $2H_2O$, and the remaining $CaSO_4$, if mixed with water, will set to a hard mass, called *plaster of Paris.* When the gypsum is heated much above the temperature mentioned, it will not afterwards set when mixed with water.

Gypsum is soluble in about 400 parts of cold water, or 1 grain in an ounce. A solution of $CaSO_4$ is used in the laboratory as a test, to show the presence of a salt of barium when mixed with one of calcium; because solution of gypsum will produce a white precipitate of $BaSO_4$ when added to a solution of any barium salt. (See p. 76.) Gypsum is insoluble in alcohol.

A white precipitate of $CaSO_4$ is produced whenever H_2SO_4 is added to a strong solution of a calcium salt (say a solution of $CaCl_2$), but *not* if the solution is so dilute as to produce less than 1 grain of $CaSO_4$ to the ounce of solution. In testing, therefore, care must be had to the strength of the solutions used.

Native $CaSO_4$ is called *anhydrite*; $CaSO_4.2H_2O$ is named *gypsum* and *selenite*; when compact and quite white, *alabaster.*

TRICALCIUM PHOSPHATE.—PHOSPHATE OF LIME.—BONE EARTH.

[$Ca_3P_2O_8$ or $Ca''_3 2(PO_4)''' = 310$.]

THIS natural product is the most common and important of the many phosphates of calcium known to exist. Its principal artificial uses are for the production of phosphorus (see *C. N.* p. 89) and as a manure for the land. A man or animal does of course eat its bones in the shape of food before it can increase them. All animal food comes directly or indirectly from the land. Land is apt to get exhausted of its supply of phosphate, unless the same, in the shape of manure, is returned to it in proportion as each successive set of crops removes the same. From the great consumption of food in towns, and loss from present sanatory arrangements of the same food as manure to the land, bones of animals as well as guano are now brought from South America and other countries to supply this loss of phosphates. They are decomposed by H_2SO_4, and converted into *superphosphate of lime* (see *C. N.* p. 89), which, being soluble in water, can easily be absorbed by the growing plant, and are especially adapted for turnips, &c.

Tricalcium phosphate occurs as a natural mineral, called *coprolite*, and, crystallised, as the mineral *apatite*. It is insoluble in water, but is soluble in solutions of the salts HCl and HNO_3, or hydric acetate, from which solutions it may be again precipitated, unchanged, by the addition of solution of ammonia.

SULPHIDES OF CALCIUM.

CALCIUM forms several sulphides, which are more or less soluble in water.

CALCIUM CHLORIDE.

[$CaCl_2 = 111$.]

THIS body is produced in the formation and production of CO_2 from marble. (See *C. N.* p. 61.) It is of great value in the laboratory, as it is far more used than any other *desiccant* (from *de*, from, and *siccus*, dry), on account of its extraordinary deliquescent properties, absorbing moisture with a rapidity that is truly wonderful.

It melts at a red heat, is very soluble in water, and may be crystallised in long crystals containing $CaCl_2.6H_2O$. Gives a red colour to flame. Is found in some mineral waters. When mixed with snow, or powdered ice, produces a powerful freezing mixture.

CALCIUM FLUORIDE.—Fluor Spar.

[$CaF_2 = 78$.]

Found as a natural mineral in Derbyshire, Cumberland, and many other places. Occurs commonly in cubical crystals. When coloured blue, forms the *blue John* of Derbyshire.

It is the chief source of all the compounds of fluorine; it is insoluble in water and most other liquids; it is decomposed by H_2SO_4, with production of hydric fluoride, HF (see *C. N.* p. 81), which is used, both as a gas and in solution in water, for writing on glass.

CHLORIDE OF LIME.—Bleaching Powder.

This substance is now manufactured in enormous quantities, owing to its great value as a disinfectant and bleaching agent, and its cheapness.

A solution of the salt HCl is now obtained as a waste product from the manufacture of carbonate of soda. (See p. 57.)

Chlorine gas is easily obtained by mixing HCl and MnO_2 together. (See *C. N.* p. 75.) The mineral manganese (MnO_2) is not expensive.

Lime can easily be obtained; and, as there is a great demand for it, and an abundant supply of it, its price is also low.

Chloride of lime is made by placing slaked lime on shelves in a room, and allowing Cl gas to pass over them. The lime absorbs the gas freely, bleaching powder being the result. This powder contains (1) a salt called calcium hypochlorite ($CaCl_2O_2$ or $Ca2ClO$); (2) calcium chloride ($CaCl_2$) formed at the same time; and (3), as the operation is never carried on so as to decompose the whole of the lime, there is also free CaO present (see *C. N.* p. 76)—

$$2CaH_2O_2 + 2Cl_2 = CaCl_2 + CaCl_2O_2 + 2H_2O$$

or

$$2(Ca2HO) + 4Cl = CaCl_2 + Ca2ClO + 2H_2O$$

The powder absorbs moisture from the air, owing to the $CaCl_2$ which it contains. It smells differently from Cl, and is partially soluble in water; the $CaCl_2$ and the $CaCl_2O_2$ dissolve, leaving the excess of lime behind. This solution is used for bleaching, as then the solid lime does not become mixed with the fabric to be bleached.

When exposed to the air, the calcium hypochlorite slowly decomposes, liberating O and Cl gases, both of which possess great bleaching and disinfecting properties, when thus liberated in the nascent state. (See p. 25.)

CALCIUM OXALATE.

[$CaC_2O_4.2H_2O$ or $CaO.C_2O_3+2H_2O = 128+36.$]

THIS body is only important as being the form in which calcium is generally precipitated from its solutions by the action of ammonium oxalate [$2(H_4N)'.(C_2O_4)''+2H_2O$]. It is a white powder, insoluble in water. It is the only oxalate which is insoluble in hydric acetate ($H\bar{A}$ or $H.C_2H_3O_2$). At 100° C. it loses its $2H_2O$, and at a *dull* red heat is converted into calcium carbonate ($CaCO_3$), the gas CO escaping—

$$CaC_2O_4 = CaCO_3 + CO$$

TESTS FOR SALTS OF LIME.

The *soluble carbonates* give white insoluble precipitates of calcium carbonate.

Dihydric sulphate (in strong solutions only) gives a white precipitate of $CaSO_4$. (See p. 82.)

Ammonium oxalate gives a white precipitate of CaC_2O_4, which is *insoluble* in hydric acetate, but soluble in hydric chloride.

Calcium chloride gives a red colour to flame, which must not be confounded with the similar colours produced by strontium and lithium.

MAGNESIUM.

[*Symbol*, Mg; *Atomic Proportion*, 24; *S.g.* 1·74.]

Occurrence.—Magnesium salts are widely distributed through the crust of the earth, but do not occur in such gigantic masses as the calcium salts. *Dolomite*, or magnesian limestone, is a double carbonate of lime and magnesia ($CaCO_3.MgCO_3$); it forms a most excellent building stone, and is often used as a source of magnesia (MgO). United with Cl, as magnesium chloride ($MgCl_2$), it is found in large quantities in sea-water. The silicates of magnesium form a very numerous and complicated class of minerals; they are found in *mica*, *hornblende*, *serpentine*, and *olivine*, united with other silicates; and in *talc*, French chalk or *soapstone*, and *meerschaum*, of which pipes are made; it occurs also, united with H_2O, as nearly pure silicate of magnesium.

History.—First obtained in very small quantities by Sir H. Davy, in 1808, and called magnium, and afterwards magnesium. Bussy, in 1830, obtained it in larger quantities by fusing magnesium chloride ($MgCl_2$) and K together.

Preparation. — By decomposing fused magnesium chloride ($MgCl_2$) by an electric current.

By fusing a mixture of powdered $MgCl_2$, NaCl, and CaF_2, with small pieces of metallic sodium (Na).

Properties.—A light, silver-white, malleable, ductile metal. Does not decompose cold water, but does decompose warm water. (See p. 9.) May be kept in dry air, as it does not very easily oxidise. When heated, it melts, and, if exposed to the air, takes fire, burning with a very brilliant white flame, which has been used to take photographs at night. Mg is now made in large quantities in Manchester, and, as wire, ribbon, and powder, it is sold to burn for photographic and other purposes. The white ash left by the burnt metal is magnesium oxide, or magnesia (MgO). The metal takes fire when thrown upon strong solution of HCl.

Magnesium with Oxygen.

There is only one known oxide of magnesium, namely, the common *magnesia* (MgO).

MAGNESIUM OXIDE.—Magnesia.

[$MgO = 40$.]

Magnesia was first recognised as a separate earth by Black, in 1755. It had been previously confounded with lime.

This body forms a white, light, bulky, tasteless, infusible, and, in water, almost insoluble powder, often known as *calcined magnesia.*

It may be obtained, like CaO, by heating the carbonate ($MgCO_3$) to redness.

When mixed with water, it gradually *slakes,* without becoming hot, forming a hydrate, MgH_2O_2, or $MgO.H_2O$, or Mg2HO. This hydrate slowly absorbs CO_2 from the air.

MAGNESIUM CARBONATE.

[$MgCO_3 = 84.$]

FOUND native in the mineral *magnesite*, and, mixed with calcium carbonate, in *dolomite*, or magnesium limestone ($MgCO_3.CaCO_3$).

PREPARATION.—This salt is difficult to prepare pure, but may be made by mixing boiling solutions of K_2CO_3 and magnesium chloride ($MgCl_2$), and, when cold, making a solution of $MgCO_3$ in CO_2 gas, allowing this solution gradually to deposit crystals of $MgCO_3.3H_2O$.

The carbonate of magnesia of the shops, known as *magnesia alba*, is a mixture of magnesium carbonate ($MgCO_3$) and magnesium hydrate (MgH_2O_2), in varying proportions, according to the method of manufacture.

The usual mode is to mix a boiling solution of magnesium sulphate with a boiling solution of sodium carbonate, when a white bulky precipitate of magnesia alba forms.

The knowledge of magnesia alba as a medicine was first promulgated from Rome, in the beginning of the 18th century.

Salts of magnesium are not precipitated by soluble carbonates in the cold; the solutions or the mixture must be boiled, when magnesia alba, and *not* pure magnesium carbonate, is precipitated.

A double salt, containing $K_2CO_3.2MgCO_3 + 9H_2O$, may also be formed.

MAGNESIUM SULPHATE.—EPSOM SALTS.

[$MgSO_4.7H_2O = 120+126.$]

THIS salt is found naturally in many spring-waters, as at Seidlitz; its name of Epsom salts is derived from the circumstance of its former preparation out of the water of certain springs in the neighbourhood of Epsom, which *naturally* contain the salt.

Magnesium carbonate is found naturally in magnesian limestone. Gypsum ($CaSO_4$) is found naturally in most spring waters. When $MgCO_3$ and $CaSO_4$ are brought together, they form $CaCO_3$ and $MgSO_4$. So a spring water containing gypsum having to filter through or permeate a bed of magnesian limestone would afterwards contain magnesium sulphate.

PREPARATION.—Dolomite ($MgCO_3.CaCO_3$) is burned in kilns; the MgO and CaO thus obtained is slaked, and afterwards well washed with water, which dissolves $CaO.H_2O$, but does not dissolve $MgO.H_2O$. The mixed earths are now decomposed by H_2SO_4, which converts them into $MgSO_4$ and $CaSO_4$. The $MgSO_4$ is very soluble in water (1 to 3), the $CaSO_4$ only soluble to the extent of 1 grain to an ounce of water (1 to 400). The strong solution of $MgSO_4$ is separated from the insoluble mud of $CaSO_4$ and evaporated.

PROPERTIES.—A salt with a nasty bitter taste, called in German bitter salt (*bittersalz*). Soluble in three times its weight of cold water. The crystals resemble those of hydric oxalate (*oxalic acid*), but on a close examination they are found to belong to a different system.

At a temperature below 300°, the crystals lose six out of the seven molecules of H_2O; the seventh is retained up to a temperature of 400°. This last atom is capable of substitution, as, for example, by potassium sulphate; thus—

$$MgSO_4+K_2SO_4.6H_2O$$

This double salt will crystallise in the same form as magnesium sulphate.

MAGNESIUM SULPHIDE

Is slightly soluble in water; may be formed by mixing magnesium sulphate and potassium sulphide.

MAGNESIUM CHLORIDE.

[$MgCl_2$.]

THIS salt is contained naturally in sea-water. It is very deliquescent, and its presence is the reason why sea-salt and many kinds of common salt run to water in damp weather. It may be obtained by decomposing the native carbonate or magnesia alba in solution of HCl.

If it is required to evaporate the salt to dryness, or fuse it, the solution must be mixed with some ammonium chloride; this prevents the partial decomposition of the $MgCl_2$ which otherwise takes place by evaporation to dryness.

PHOSPHATES OF MAGNESIUM.

A NUMBER of phosphates containing Mg may be formed; the only one, however, of importance is the ammonio-magnesium phosphate, often called *triple phosphate* $(H_4N)'Mg''(PO_4)'''.6H_2O$. In analysis Mg is generally precipitated as this salt, because it is quite insoluble in water containing a little solution of ammonia. It may be formed by adding to a solution of a Mg salt, firstly, some solution of ammonium chloride, and, secondly, some solution of ammonia. At this stage there should be no precipitate formed. The chloride is added to keep the Mg salt in solution, because solution of ammonia precipitates Mg salts in its absence. Solution of common phosphate of soda may now be added, when, upon briskly stirring the mixture, a copious *crystalline* precipitate of triple phosphate is thrown down.

When $2[(H_4N)Mg(PO_4).6H_2O]$ is gently heated, first the $12H_2O$ is driven off, and at a higher temperature the ammonia, leaving magnesium pyrophosphate ($Mg_2P_2O_7$); in analysis, this may be weighed, and the quantity of Mg calculated from it.

TESTS FOR SALTS OF MAGNESIUM.

THE *soluble carbonates* give a white precipitate of magnesia alba, when boiled with salts of Mg.

Solution of ammonia gives a slight precipitate in Mg salts, in the absence of ammonium chloride, but no precipitate if it be present in the solution.

Ammonium phosphate gives a white *crystalline* precipitate of ammonio-magnesium phosphate. (See p. 90.)

If moistened with a solution of cobalt nitrate, and heated before the blowpipe, salts of Mg give a pink colour.

GROUP C.—METALS OF THE EARTHS.

THE earths are insoluble in water, and have neither taste nor action on test paper.

The metals of the earths are ten in number, namely—

ALUMINIUM	Al	Erbium	Er
Glucinum	G	Terbium	T
Zirconium	Zr	Cerium	Ce
Thorinum	Th	Lanthanum	La
Yttrium	Y	Didymium	D

The metal and salts of aluminium are the only bodies of any use and importance in this group.

1st. Aluminium does not form a carbonate. The soluble carbonates precipitate it as a white sesquioxide.

2nd. It is also precipitated as white sesquioxide by hydro-ammonium sulphide.

ALUMINIUM.

[*Symbol*, Al‴; *Atomic Proportion*, 27·5; *S.g.* 2·6.]

Occurrence.—As alumina (Al_2O_3), it is next to silica (SiO_2), the most abundant constituent of the earth. It is a chief component of clay, field earth, and many kinds of rock, as slate and granite; also of porcelain, earthenware, and bricks, united generally with silica; the colour of the many minerals containing it being due to small and variable quantities of Fe and other substances.

History.—This metal was first obtained by Wöhler, in 1828, by heating K in a tube, and passing aluminium chloride (Al_2Cl_6) over it. The metal K was sold at the rate of 1,000*l.* per pound weight when this discovery was made, the discoverer who had to purchase it being in the receipt of 50*l.* a year salary at the time.

Alum, the salt from which the term aluminium is derived, has been known from very early times, and until the 15th century was brought to Europe from the East. In 1728 the existence of alumina in clay was first shown. In 1826 Oersted made Al_2Cl_6, from which, two years later, the metal was obtained.

Preparation.—By passing vapour of aluminium chloride (Al_2Cl_6) over melted Na.

By fusing Na with the mineral *cryolite* ($3NaF.AlF_3$), mixed with NaCl and KCl to act as a flux. The Na combines with F, and liberates Al.

By the decomposition of the double chloride of sodium and aluminium ($2NaCl.Al_2Cl_6$) by a powerful galvanic current.

Properties.—A light, malleable and ductile metal, of the colour of zinc. Melts easily; does not easily oxidise in the air. In powder, decomposes boiling water. Is very sonorous. Forms an alloy with copper, in colour like gold, called aluminium bronze; it will not alloy with lead. It does not form an amalgam with mercury. Is now manufactured on a large scale at Newcastle-on-Tyne.

It is not dissolved by HNO_3, but is easily acted on by HCl, a chloride of the metal being formed and dissolved. The alkalies easily oxidise and dissolve it, giving off H gas.

ALUMINIUM WITH OXYGEN.

ALUMINIUM only forms one compound with O—namely, the sesquioxide, called *alumina* (Al_2O_3).

ALUMINIUM SESQUIOXIDE.—ALUMINA.

[$Al_2O_3 = 103$.]

THIS body is found native as the mineral *corundum*. Coloured by a salt of chromium, it is found as the *sapphire* and *ruby*; in a more common and impure form, as *emery*. All these minerals are noted for their extreme hardness.

PREPARATION.—By precipitating alum or any salt of aluminium by K_2CO_3 or hydro-ammonium sulphide. The precipitate is very bulky and gelatinous, and requires much washing with water to get rid of the last traces of the precipitating solution.

PROPERTIES.—White, without smell, taste, or action on test paper, and can only be melted in the oxy-hydrogen gas flame. When freshly precipitated, it is a hydrate, $Al_2O_3.3H_2O$; is easily decomposed and dissolved as chloride by HCl, and as potassium aluminate (K_3AlO_3) by KHO; but after drying and roasting, it is very difficult of solution by any means, unless it is melted first with KHO.

This oxide is peculiar, as it possesses the properties of a non-basylous radical at one time, and of a basylous radical at others: thus there is a native mineral, called *spinell*, which is magnesium aluminate ($MgAl_2O_4$), and also a zinc aluminate ($ZnAl_2O_4$).

Sodium hydrate (NaHO) and solution of ammonia act in the same way as KHO—that is, they first precipitate hydrate of alumina, and then redissolve it.

Alumina has the peculiar property of uniting itself to certain vegetable substances, and this property of alumina has been taken advantage of by the dyer and colour-maker. When to a solution of alum and cochineal K_2CO_3 is added, the alumina is precipitated, not as a *white* hydrate, but coloured by the cochineal, as the colour is precipitated with the alumina, forming a *lake*. As a *mordant* (from *mordeo*, I bite), it is used to fix colours on calico. The cloth is soaked in a preparation of alumina, and then dipped in a solution of the dye; the dye thus becomes fixed to the alumina and the cloth, and cannot be washed out. *Alum* is the salt ordinarily used to supply pure alumina for the dyer, and is extensively manufactured for this purpose.

ALUMINIUM SESQUISULPHATE.

[$Al'''_2 3(SO_4)''.18H_2O$.]

THIS salt may be formed by dissolving alumina in H_2SO_4; on the large scale, clay free from Fe is heated with H_2SO_4, and the mass dissolved in water.

ALUM.—DOUBLE SULPHATE OF ALUMINIUM AND POTASSIUM.

[$KAl\,2SO_4.12H_2O$ or $K_2SO_4.Al_2 3SO_4.24H_2O$.]

PREPARATION.—1st. By lixiviating certain volcanic earths which contain it.

2nd. By burning and lixiviating alum-stone.

In some parts of the earth, as in Civita Vecchia, near Rome, a rock is found, called *alum-stone*, which contains *all the* elements necessary for the production of alum.

Portions of the rock are gently burned in kilns, and lixiviated in water, when alum, called Roman alum, is obtained.

3rd. By burning and lixiviating certain shales, or alum *schists*, and adding a potassium salt to the solution.

In England at Whitby, and at Glasgow, a bituminous *shale*, which contains iron disulphide (*pyrites*) (FeS_2) and aluminium silicate, is slowly roasted. During the roasting, the S of the pyrites oxidises, and forms, with the H_2O contained in the shale, H_2SO_4; this decomposes some of the aluminium silicate. The mass is afterwards lixiviated, when a solution of iron sulphate ($FeSO_4$) and aluminium sulphate (Al_23SO_4) is obtained. Potassium chloride (KCl_2) is now added to the mixed sulphates, when $FeCl_2$ is formed and K_2SO_4; this latter crystallises out of the solution, in company with the aluminium sulphate, as *alum*.

PROPERTIES.—Crystallises in large colourless *octahedra*, soluble in 18 parts of cold water; the solution has a very astringent taste (from *ad*, to, and *stringo*, I tie fast), drawing the sides of the mouth together. When heated, it loses its water of crystallisation, and then forms *burnt alum*.

It is largely used in dyeing and the arts; also in medicine, as a gargle for the throat.

Alum may be viewed as a *double salt*. (See p. 31.)

ISOMORPHISM.—The peculiarity of two bodies crystallising in the same form, although of different chemical composition, is well illustrated by the alums.

Chrome alum, made from Cr_2O_3, instead of Al_2O_3, has the following formula:—$KCr2SO_4.12H_2O$, that is, the same formula as common alum, Cr being substituted for Al. Its solution is dark green, and its crystals a dark red; it crystallises, however, in *octahedra* like common alum.

Iron alum, made from Fe_2O_3, instead of Al_2O_3, crystallises in violet *octahedra* with the formula $KFe2SO_4.12H_2O$ —that is, an alum with Fe, instead of Al:—

Common alum . .	$KAl\ 2SO_4.12H_2O$
Chrome „ . .	$KCr\ 2SO_4.12H_2O$
Iron „ . .	$KFe\ 2SO_4.12H_2O$

If a small crystal of chrome alum be placed in a solution of common alum, the crystal will be increased, by the deposition of a coating of colourless common alum upon the chrome alum.

If *common, chrome,* and *iron* alum be mixed together, crystals will be formed, not of each alum, but containing all these alums, in the proportion in which the salts were mixed.

If an octahedron of wood or glass be placed in an alum solution, instead of an alum crystal, the salt will not deposit evenly upon it, but in small octahedral crystals on the surface of the foreign substance.

If one surface of a crystal of alum be covered with a thin coating of varnish, the crystal will not grow in that direction; it will deposit on all the surfaces but the varnished one.

Because Al_2O_3, Cr_2O_3, Fe_2O_3 can replace one another in such a class as the alums, they are said to be isomorphous.

Further, K_2O may be replaced in an alum by either Na_2O, forming a soda alum, or $(H_4N)_2O$, forming an ammonia alum, because K_2O, Na_2O, $(H_4N)_2O$ are all isomorphous with one another (see *C. N.* p. 29):—

Potassium alum . .	$KAl\ 2SO_4.12H_2O$
Sodium „ . .	$NaAl\ 2SO_4.12H_2O$
Ammonium „ . .	$(H_4N)Al\ 2SO_4.12H_2O$

ALUMINIUM CHLORIDE.

$[Al_2Cl_6 = 268.]$

THIS is the salt from which the metal Al is generally made.

Preparation.—By passing dry Cl gas over a mixture of Al_2O_3 and C, heated to redness. (See p. 13.) The Al_2Cl_6 distils over, and must be condensed, while CO escapes.

Properties.—A yellow, crystalline, deliquescent solid, which easily melts and sublimes. Rapidly dissolves in water, with a hissing sound.

If Al_2O_3 be decomposed by 6HCl, the solution of Al_2Cl_6 or $2(AlCl_3)$ formed, decomposes when evaporated to dryness.

The salt is now made on the large scale.

ALUMINIUM PHOSPHATES.

The blue *turquoise* is a native hydrated phosphate of aluminium, and the *lazulite* is a double phosphate. A white gelatinous hydrated phosphate may be formed by mixing sodium phosphate with alum, which is easily soluble in HCl.

ALUMINIUM SILICATES.

There are a great many silicates of aluminium, the most common and useful of which is clay ($Al_2O_3 2SiO_2 . 2H_2O$), or hydrated silicate of aluminium. It is principally derived from the mineral *felspar*, which is a double silicate of aluminium and potassium, or sometimes sodium, when the mineral is called *albite*; the composition of the felspars may be shown by the following formula, where M′ represents any univalent metallic basylous radical:—

$$M'_2O.3SiO_2.Al_2O_3.3SiO_2$$

Pure clay is white, and, when mixed with water, forms a stiff, tenacious mass, of a highly *plastic* nature (from πλάσσω, I form, mould), on which account it is used to make bricks, tiles, earthenware, pottery, porcelain, &c.

Fire-clay.—When clay is pure, it is almost infusible, while ordinary red clay, of which common bricks are made, contains iron, lime, and other impurities, and is consequently always more or less fusible in the furnace heat. Fire-clay, therefore, of which fire-bricks and crucibles are made, is a purer kind of clay, that does not fuse in the fire.

KAOLIN is a very white and pure clay, used for making porcelain.

PIPECLAY is a white clay, found in Dorsetshire, and is used for making tobacco-pipes. It fuses at an intense heat.

FULLER'S EARTH is a porous kind of clay, found near Reigate; it has a peculiar power to absorb oily substances, and is used to remove grease, &c. from cloth.

PORCELAIN AND POTTERY WARE.—After clay has been moulded into shape, and during the process of baking, it shrinks very considerably in size, and not always regularly, so that the article generally loses its perfect shape. It is the object of the manufacturer of earthenware to make such a mixture as shall retain its shape while baking, be sufficiently close and dense to bear constant handling, and not break too easily. This is accomplished by taking *clay*, which gives plasticity and enables the mass to be moulded into shape, *chalk*, *sand*, and felspar, called by the Chinese potters *petuntze*, all which together in proper proportion, partially melt in the fire, giving a half-glassy appearance to good porcelain. The different kinds of earthenware are caused by the varied purity and proportion of the ingredients. The degree of heat also to which the manufactured ware is subjected has much to do with the article made.

GLAZE is a composition, usually liquid, with which the surface of any article is covered, and which shall fuse on the surface, and become fixed when heated in the fire. Salt, borax, and lead oxide, &c., are used in glazing.

TESTS FOR SALTS OF ALUMINIUM.

When moistened with solution of cobalt, and heated before the blowpipe, they become of a beautiful blue colour.

Hydrates of K *and* Na first precipitate alumina, but redissolve the precipitate if more hydrate be added than is required for the precipitation.

Solution of ammonia precipitates alumina, and redissolves small quantities of the precipitate when added in excess.

The soluble carbonates precipitate alumina, and do not redissolve the precipitate if added in excess.

There is no carbonate of aluminium.

Hydro-ammonium sulphide precipitates alumina, and *not a sulphide.*

SECOND DIVISION.—HEAVY METALS.

Group A.—Iron Group.

The metals of this group tarnish or *rust* in moist air; decompose water at a red heat only; and are not precipitated from their solutions by H_2S in presence of HCl (see pp. 9 and 11):—

Zinc	Zn	Manganese	Mn
Iron	Fe	Cobalt	Co
Chromium	Cr	Nickel	Ni

These metals are usually considered to be divalent. They all have a s.g. over 5, and are hard. Their oxides, carbonates, and sulphides are insoluble in water.

ZINC.

[*Symbol,* Zn; *Atomic Proportion,* 65; *S.g.* 7·00.]

Occurrence.—Found native as a carbonate, called *calamine* or *galmei,* and as a sulphide, called *blende.* It is

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never met with native. There is no practical method of obtaining Zn from its silicate.

HISTORY.—Zinc, or, as it is frequently called in commerce, *spelter*, has been known from the time of Paracelsus. The ore of zinc called calamine was used by the ancient Greeks in the manufacture of brass.

PREPARATION.—The ores of zinc are first crushed and roasted; after which they are distilled with powdered coal or coke. The crucibles used to distil it are covered and provided with an iron exit tube, passing through the bottom of the crucible, down which the Zn distils (distillation *per descensum*), or it is distilled from small earthen retorts.

PROPERTIES.—Pure zinc is a hardish, blue-white metal, with considerable lustre; it is brittle in the cold, but can be rolled into thin sheets and drawn into wire at a temperature of about boiling water.

It forms several alloys, of which *brass*, composed of copper and zinc, is the most important.

Galvanised iron is made by dipping clean sheet iron in a bath of melted Zn.

Zn is easily dissolved by HCl and H_2SO_4, liberating H; it is also dissolved by KHO, with liberation of H and formation of K_2ZnO_2.

In testing for arsenic, pure zinc is required; the zinc of commerce is never pure, and H liberated by it always possesses a disagreeable smell, and contains traces of arsenic (As). (Marsh's Test.)

Zinc is a very basylous radical, and precipitates many of the metals from their solution, becoming itself dissolved in their stead.

ZINC WITH OXYGEN.

THERE *is only one well ascertained compound of Zn with O*—namely, ZINC OXIDE (ZnO).

ZINC OXIDE.

[$ZnO = 81$.]

Zinc burns in the air, when sufficiently heated, with a bluish flame, forming ZnO, which is yellow when hot, but white when cold.

The *hydrated oxide* ($ZnO.H_2O$) may be precipitated from solution of zinc salt by the addition of K or Na hydrate, care being taken not to add too much of the precipitant, as the zinc hydrate is soluble in K hydrate.

As *zinc white*, the oxide is used for a paint, because it is not blackened by H_2S, like the paint called *white lead.*

ZINC CARBONATE.—Calamine.

[$ZnCO_3 = 125$.]

Occurs native as *calamine.*

The white precipitate formed by the addition of the soluble carbonates to solution of a salt of zinc is not a true carbonate, but a mixture of carbonate and oxide of zinc. *This white precipitate is soluble in solution of ammonium carbonate,* but not in K or Na carbonate.

ZINC CHLORIDE.—Burnett's Fluid.

[$ZnCl_2 = 136$.]

This salt is obtained by dissolving Zn in 2HCl; the solution of $ZnCl_2$ formed may be evaporated, when the salt is obtained in the solid form.

This solid salt is very deliquescent and corrosive to the skin; its solution is used as a disinfectant, called *Sir William Burnett's Disinfecting Fluid.* It removes foul odours from decomposing matter when mixed with them. It also preserves wood and fibre from decay; it is therefore an *antiseptic.*

In solution it is used as a flux by the zinc solderers.

ZINC SULPHIDE.—BLENDE.

[$ZnS = 97.$]

FOUND native, as *blende*, in dark red masses or crystals, which are decomposed by HNO_3 or a mixture of HNO_3 and HCl.

A *hydrated* sulphide of zinc is thrown down from a solution of a salt of Zn by hydro-ammonium sulphide, as a white gelatinous precipitate; but not by H_2S in the presence of HCl. The acetate is precipitated, however, by H_2S in presence of hydric acetate.

ZINC SULPHATE.—WHITE VITRIOL.

[$ZnSO_4.7H_2O = 161+126.$]

THIS salt is formed in the preparation of H from Zn and H_2SO_4. (See *C. N.* p. 52.) It may also be prepared by roasting the sulphide at a *low* temperature. (See pp. 13, 95.) The compound obtained by roasting a *sulphide* depends greatly upon the circumstances under which it is roasted; thus, if roasted at a gentle heat it is partly converted into a *sulphate*, while at a higher temperature an *oxide* is formed.

Zinc sulphate, or *white vitriol*, effloresces in dry air; it is soluble in 2½ parts of cold water. It loses its water of crystallisation by heat, and forms double salts, like magnesium sulphate. It is used in medicine as an emetic (from ἐμέω, to vomit).

TESTS FOR THE SALTS OF ZINC.

THE soluble salts of zinc are poisonous in large doses; in smaller quantities they act powerfully and rapidly as emetics.

The alkalies give a white precipitate of hydrated oxide, soluble in excess of alkali.

The soluble carbonates give a white precipitate, *soluble in ammonium carbonate.*

Hydro-ammonium sulphide gives a white precipitate of sulphide of zinc; H_2S gives no precipitate in presence of HCl.

If moistened with solution of cobalt, and heated in the blowpipe flame, compounds of Zn give a *green* infusible residue.

IRON.

[*Symbol,* Fe (Ferrum); *Atomic Proportion,* 56; *S.g.* 7·8.]

OCCURRENCE.—Iron, as oxide, is one of the most widely distributed of the elements. It occurs native in meteoric stones, but very seldom otherwise. As iron sesquioxide (Fe_2O_3), it gives the red colour to clays and sandstone, and is found in the blood of animals.

The ores from which iron is obtained are numerous; the chief are: *hæmatite,* red (Fe_2O_3) and brown ($Fe_2O_3.3H_2O$); *clay ironstone,* containing iron carbonate, clay, and often lime, from which large quantities of iron are obtained in England; *magnetic iron ore,* or *loadstone* ($FeO.Fe_2O_3$); *specular* iron ore similar to red hæmatite, and *spathic* iron ore which is pure iron carbonate.

HISTORY.—Iron (♂) or Mars was one of the seven metals known to the ancients. It has been reduced from its ores and used by some of the natives of Africa and India from præhistoric times.

PREPARATION.—The production of iron is one of the most important manufacturing operations, and is very extensively carried on in England, Belgium, and many other countries.

The ore is frequently first roasted, to get rid of water, CO_2, S, &c. Several ores are then mixed together, those containing much iron, rich ores, with some containing much less iron, poor ores.

The next process is to heat the roasted ore, which always contains the metal in the form of oxide, to intense whiteness in a large furnace, from 50 to 80 feet high, called a *blast* or *smelting* furnace. Coal and coke are employed to produce the necessary heat, and calcium carbonate, *chalk* ($CaCO_3$), is added to form a flux with the clay, ashes of coal, and other impurities. The heat of the fire is kept up by means of large blowers worked by machinery. The bottom of the furnace is closed in with fire-bricks, small holes for the *air-pipes*, called *tuyères*, and for the iron and slag to run out of, alone being left.

In the furnace the heated C takes O from the iron oxide (*the ore*), forming CO; the Fe, as soon as separated, is melted, owing to the intense heat, and, being liquid and heavy, runs to the bottom of the furnace, where it collects. The $CaCO_3$ is soon converted in the fire to CaO, and, when sufficiently heated, unites with the clay, sand, and other impurities, forming a kind of glass or fusible porcelain, called *slag*, which, being liquid at the intense heat of the furnace, runs down, and, being light, swims on the surface of the melted iron. (See p. 16.) Long flames issue from the mouths of these smelting furnaces, where the CO and other gases are completely burned in the air.

The last process consists in opening certain small holes, which are stopped by plugs of clay, and allowing the liquid iron to run out into channels made in sand, where it cools; each channel of iron when cold, is called a *pig*; hence the name *pig-iron*. The slag is also run out, from a small hole in the furnace, whenever the workman thinks it necessary, which he knows from practice. Fresh supplies of (1) roasted ore, (2) coal, and (3) chalk are being continually thrown in at the top of the furnace, as its contents are burnt away or removed, while (4) air is supplied at the bottom. If the air is blown in cold, it forms the *cold* blast, but if it be heated before it is blown into the furnace, it forms the *hot* blast.

Pig-iron, or *cast iron*, as obtained from the blast furnace, is not pure iron, and requires to be further operated upon before it is fit for use; it contains variable quantities of C, Si, As, S, P, Mn, Al, Ca, &c. Cast iron is of various kinds, as *white*, *mottled*, *grey*, &c., according to variations in the circumstances of manufacture, but *all cast iron is especially remarkable for the amount of* C *it contains* (3 to 5 per cent.). Cast iron is hard, brittle, and cannot be welded; neither is it elastic. It corresponds to the formula Fe_4C (carbide of iron).

Wrought iron is made from cast iron by burning away a great deal of the C, Si, and other impurities in a furnace, called a *refinery*, where air is allowed to enter to the melted iron, and a coating of iron oxide and SiO_2 is formed as a kind of slag. It is next placed in a *puddling* furnace, where it is again heated, covered by scales of iron oxide; or sometimes the *pigs* are at once placed in the *puddling* furnace. The puddling furnace is a kind of reverberatory furnace. Here nearly all the C and Si are oxidised; and after the iron has been in the furnace for about two hours, it is taken in lumps, and, while hot, tremendously hammered or squeezed between rollers, to get rid of all slag, &c., and is now *wrought iron*. *Wrought iron contains very little* C ($\frac{1}{2}$ per cent.); it is soft, may be bent without breaking, tough, and is capable of being welded. It cannot be rolled into very thin leaves, but forms thin and remarkably tough wire. It is far more difficult to melt than cast iron. It is magnetic.

With very pure iron ore, iron equal to wrought iron may be made at once from the ore by using charcoal instead of coal.

Steel is made by taking bars of wrought iron, and heating them for many hours, covered in charcoal powder, by which process the iron takes up some of the C, and becomes possessed of the properties of *steel*. Steel *therefore contains more* C *than wrought iron, but does not contain* *so much as cast iron.* Steel is grey-white, and may be made either hard enough to scratch glass, or soft enough to be cut with a file; if heated and quickly cooled by dipping in water, it is hard, but if allowed to cool slowly, it is soft, and the hardness depends upon the temperature to which the steel was heated before being suddenly cooled. Steel is elastic, fine-grained, fusible, and capable of a very high polish. It retains magnetism, and so is used to make permanent magnets.

PROPERTIES.—Iron is soluble in HCl and H_2SO_4, forming green solutions of $Fe''Cl'_2$ and $Fe''(SO_4)''$, with liberation of H. It is also soluble in a solution of HNO_3, of a s.g. not higher than 1·3; in a very strong solution of HNO_3, iron may be kept for years without being dissolved (*the passive condition of iron*).

It precipitates many metals from their solutions, as copper (Cu), being itself dissolved, taking the place of the precipitated metal. It burns with brilliancy in O gas, if first heated red-hot. At a red heat, it decomposes water, forming an oxide, the same compound as the black scales from the smith's anvil—namely, the magnetic oxide of iron.

Iron does not oxidise or rust in dry, but very quickly in moist air, forming a red hydrated sesquioxide.

IRON WITH OXYGEN.

THERE are four compounds of iron with O—

White Ferrous oxide, or Iron oxide .	$Fe''O''$
Red Ferric oxide, or Iron sesquioxide .	$Fe'''_2O''_3$
Black magnetic oxide of iron . . .	$Fe''O''.Fe'''_2O''_3$
Iron trioxide	$Fe^{vi}O''_3$

All the salts of iron which are commonly met with may be divided into two distinct and separate classes—namely, those which contain Fe'' as a divalent radical, as $Fe''O''$, and those which contain Fe''' as a trivalent radical, as $Fe'''_2O''_3$. The salts of the former are generally green, and form a carbonate ($FeCO_3$), while those of the latter are red in colour, and, like all the sesquioxides, does not form a carbonate.

FERROUS OXIDE.—Iron Oxide.

[$Fe''O'' = 72$.]

This salt is very difficult to separate from combination and obtain pure, owing to the extraordinary affinity it possesses for O. It is white when freshly precipitated, but rapidly passes to grey, green, blue, mud colour, to sesquioxide. It may be formed by precipitating iron sulphate [$Fe''(SO_4)''$] by K_2CO_3.

FERRIC OXIDE.—Iron Sesquioxide.

[$Fe'''_2O''_3 = 160$.]

Found native in the minerals hæmatite and specular iron ore. It also forms the *bloodstone.* Made artificially by heating $Fe''SO_4$ in the air, when a red powder is formed, called *crocus of Mars*, and used to polish glass, jewellery, &c.

Hydrate of Iron Sesquioxide; Ferric Hydrate.—Occurs native as *brown* hæmatite. It may be thrown down as a red-brown precipitate by adding solution of ammonia or its carbonate to solution of iron sesquichloride ($Fe'''_2Cl'_6$), and well washing the precipitate. It is bulky, flocculent, and easily soluble in HCl. It forms several *hydrates*, as $Fe_2O_3.H_2O$ and $Fe_2O_3.3H_2O$; they all lose their H_2O by heat.

MAGNETIC OXIDE OF IRON.

[$Fe''O''.Fe'''_2O''_3$ or $Fe^{2\frac{2}{3}}_3O''_4$.]

Occurs native as *loadstone,* and composes the chief part of the black scales which fall from the smith's anvil when wrought iron is worked. It does not form any distinct class of salts.

FERRIC OR IRON TRIOXIDE.

$[Fe^{vi}O_3.]$

HAS never been isolated, but may be obtained as ferrate of potash.

FERROUS OR IRON CARBONATE.

$[Fe''CO_3'' = 116.]$

THIS body is found native in very large quantities, forming a most valuable ore of iron, which, when pure, often in fine rhombic crystals, is called *spathic* iron ore, and, when mixed with clay, forms the *clay* iron ore.

It may be produced by adding hot solution of a soluble carbonate to a hot solution of iron sulphate $[Fe''(SO_4)'']$. It is precipitated white, but rapidly turns green, and eventually red, by absorption of O from the air. It is soluble in CO_2, and is frequently found dissolved in spring water, to which it imparts an inky taste; such water, when exposed to the air, gradually loses CO_2, the iron carbonate ($FeCO_3$) is deposited; which, however, soon absorbs O, and turns to red sesquioxide ($Fe'''_2O''_3$), and is the cause of the red deposit by 'chalybeate' (from χάλυψ, steel) or ferruginous springs.

FERROUS OR IRON SULPHATE.—COPPERAS OR GREEN VITRIOL.

$[Fe''SO_4''+7H_2O = 152+126.]$

THIS salt is one of the most common, important, and easily made iron salts of the divalent class. Iron sulphide, or old scraps of iron, are put into dilute dihydric sulphate $[H'_2(SO_4)'']$, when H_2 is given off, and $Fe''(SO_4)''$ is formed and dissolved in the liquid, which is then filtered and left to crystallise. (See *C. N.* pp. 72, 87.) Green crystals, soluble in water; both the crystals and solution gradually absorb O from the air, red salts of the trivalent class being formed.

It is used for making ink and a black dye, as, when mixed with *tannin*, it gives a black precipitate of tannate of iron. Also, when mixed with HNO_3, it turns dark olive coloured (see *C. N.* p. 68), owing to the formation of nitric oxide, which dissolves in the solution of sulphate.

FERRIC SULPHATE OR IRON SESQUISULPHATE.

$[Fe'''_2 3(SO_4)'' = 400.]$

Is made by dissolving $Fe'''_2 O''_3$ in H_2SO_4, or by adding the necessary quantity of H_2SO_4 to convert $Fe''(SO_4)''$ into $Fe_2 3SO_4$, and oxidising the Fe'' to Fe''' by HNO_3. It does not crystallise, but may be obtained as a yellowish-white mass, and forms a dark red solution.

SULPHIDES OF IRON.

Iron unites with S in several proportions, only two of which are of much importance, namely—

Black Ferrous or Iron sulphide . .	$Fe''S''$
Yellow Ferric or Iron disulphide Iron pyrites .	$Fe^{iv}S''_2$

FERROUS OR IRON SULPHIDE.

$[Fe''S'' = 88.]$

Preparation.—By heating a bar of iron white-hot, and rubbing it with a stick of S, when liquid FeS runs down, of a reddish-brown colour.

2. By throwing a mixture of iron filings and S into a red-hot crucible.

3. As a hydrate, by precipitating $FeSO_4$ by hydro-ammonium sulphide.

Properties.—Black, decomposed by HCl and H_2SO_4, with liberation of H_2S. Insoluble in water and the alkalies.

FERRIC OR IRON DISULPHIDE.—IRON PYRITES.

[FeS_2.]

FOUND *native* in *bright yellow crystals*, with a metallic lustre. (See p. 1.) Not acted upon by dilute solutions of H salts, but easily decomposed by moderately strong solution of HNO_3. When heated sufficiently in the air, its S takes fire and burns, giving off SO_2, on which account it is used in the manufacture of H_2SO_4. (See *C. N.* pp. 82, 84.)

NITRATES OF IRON.

IF iron is placed in HNO_3, of s.g. 1·2, it rapidly dissolves, red fumes are given off, and $Fe'''3(NO_3)'$ formed. If FeS is dissolved in cold and very dilute solution of HNO_3, a compound $Fe''2(NO_3)'$ is formed. If iron is placed in very strong HNO_3, it remains bright and unacted upon.

IRON WITH CHLORINE.

Two important compounds of Fe and Cl are known—

Green Ferrous or Iron chloride . . .	$Fe''Cl'_2$
Yellow or red Ferric chloride or Iron sesquichloride . . .	$Fe'''Cl'_3$ or $Fe'''_2Cl'_6$

The former is made either by passing HCl over red-hot metallic iron, or by dissolving Fe in HCl, when a green solution is obtained of $FeCl_2$.

The latter may be made either by passing Cl over red-hot metallic iron, or by dissolving the metal in a mixture of HCl and HNO_3, when Fe_2Cl_6 or $2(FeCl_3)$ is formed.

Both the anhydrous chlorides are volatile, and may, by sublimation, be obtained in crystals; the hydrates may also be crystallised from their solutions. A solution of $Fe'''_2Cl'_6$ is always kept in the laboratory as a test.

THE Bromides, Iodides (see p. 50), and Cyanides of iron may also be formed; they are, however, of little consequence, but the double cyanides of iron are very important, namely—

POTASSIUM FERROCYANIDE.—YELLOW PRUSSIATE.

$[K_4FeCy_6+3H_2O.]$

SEE p. 51.

POTASSIUM FERRICYANIDE.—RED PRUSSIATE.

$[K_3FeCy_6.]$

THIS salt is formed by passing Cl through a solution of yellow prussiate, by which means K is removed as KCl. The salt crystallises in long red crystals.

Prussian Blue.—Discovered in 1704 in Berlin. Whenever a solution of ferrocyanide of potassium is added to a salt containing Fe''' in solution, as $Fe'''_2Cl'_6$, a beautiful blue colour is produced, called *Prussian blue*—

$$3(K_4.FeCy_6) + 2(Fe'''_2Cl'_6) = Fe_7Cy_{18} + 12KCl$$

Prussian blue (Fe_7Cy_{18}) is instantly destroyed by the alkalies, Fe_2O_3 being precipitated, and, if KHO is the alkali used, potassium ferrocyanide produced—

$$Fe_7Cy_{18} + 12KHO = 3(K_4.FeCy_6) + 2(Fe'''_2O''_3) + 6H_2O$$

The ferricyanide, or red prussiate, gives a blue precipitate with any salt of iron containing Fe'', as $Fe''SO''_4$ or $Fe''Cl'_2$, but not with salts containing Fe'''; this blue colour is called *Turnbull's blue.* It is formed as follows:—

$$2(K_3.FeCy_6) + 3(Fe''Cl'_2) = Fe_5Cy_{12} + 6KCl$$

IRON OXALATE.

$[Fe''.C_2O_4.]$

Is formed as a bright yellow precipitate, when ammonium oxalate is added to $Fe''SO_4$. If heated in a current of H, it is reduced to a black spongy mass of metallic iron, which takes fire the instant it is exposed to the air (see p. 3).

TESTS FOR THE SALTS OF IRON.

Ferrous Salts$=Fe''$.

They are not precipitated by H_2S, unless the solution is perfectly free from H salts; are precipitated black by *hydro-ammonium sulphide*. As the black sulphide precipitated is soluble in HCl and H_2SO_4, no precipitate is produced if these salts are present.

They give a blue precipitate with *potassium ferricyanide*. Soluble *carbonates* give a bulky precipitate of iron carbonate, which is white when the salts are pure, but commonly green.

The *alkalies* precipitate hydrated oxide, of a white or green colour, according to the purity of the salt precipitated.

Ferric Salts$=Fe'''$.

They produce a *white* precipitate with H_2S, which is pure S in a fine state of division, the salt, by the action of the H_2S, being reduced from the state of Fe''' to that of Fe''. *Hydro-ammonium sulphide* produces a black precipitate of iron sulphide ($Fe''S''$).

They give a blue precipitate with *potassium ferrocyanide*, best in presence of H salts, not with alkalies.

They give a blood-red colour with *potassium sulphocyanide*, best in presence of H salts, not with alkalies.

They give a black (ink) with *tannin* or tincture of galls.

Carbonates and the alkalies precipitate red hydrated oxide.

CHROMIUM.

[*Symbol*, Cr; *Atomic Proportion*, 52·5; *S.g.* 6·8.]

Occurrence.—This metal, compared with Fe, is sparingly distributed through the earth. Its most important ore is the *chrome ironstone* $Fe''O''.Cr'''_2O''_3$, chiefly obtained from Sweden and North America; it is occasionally found as lead chromate ($PbCrO_4$).

HISTORY, &c.—Obtained by Vauquelin in 1797 from the native lead chromate. The *metal* has never been used in the arts; it can be reduced from the oxide by C at a white heat. It is steel grey, hard, brittle and soluble in solution of HCl with liberation of H.

CHROMIUM WITH OXYGEN.

THERE are four compounds of chromium with O corresponding to the four oxides of iron:—

Brown Chromous or Chromium oxide .	$Cr''O''$
Green Chromic or Chromium sesquioxide .	$Cr'''_2O''_3$
An oxide corresponding to the magnetic oxide of iron	$Cr''O''.Cr'''_2O''_3$
Chromic trioxide	$Cr^{vi}O''_3$

The salts of chromium are all coloured, and from their great variety of hue the name chromium is derived (from χρῶμα, colour).

The *chromous* or *chromium oxide* is only known as a brown hydrate, which possesses even a greater affinity for O than the corresponding iron compound.

The *chromic oxide* or *chromium sesquioxide* is green and valuable as a paint. It cannot be melted, and forms hard crystals. It resembles sesquioxide of aluminium in being soluble in HCl when freshly precipitated; from this solution it may be again precipitated by the alkalies, in excess of which it is also soluble. It differs from alumina, the chrome being precipitated from its solution in KHO by boiling. Its salts form blue, green, or violet-coloured solutions. It is the colouring principle of the emerald and ruby.

CHROMIC OR CHROMIUM TRIOXIDE.

$[Cr^{vi}O''_3.]$

Is obtained in beautiful crimson crystals by adding H_2SO_4 to a solution of potassium dichromate. They are very soluble in water, easily decomposed by alcohol, vapour of ammonia, or heat, into green sesquioxide.

CHROMATES OF POTASSIUM.

THERE are two important chromates, a yellow or simple chromate (K_2CrO_4), and a red dichromate ($K_2Cr_2O_7$ or $K_2CrO_4.CrO_3$). These salts are of great importance, as they are employed to produce all the salts of chromium used as colours and paints.

The dichromate is made in large quantities from the chrome ironstone, by fusing it with a mixture KHO and KNO_3 in a reverberatory furnace; the mass is lixiviated in water, and the excess of alkali removed by HCl. The salt is then separated from the precipitated silica, &c., and allowed to crystallise.

Lead chromate is yellow, and much used as a paint. There are two chlorides of chromium.

TESTS FOR CHROMIUM SALTS.

THEY give a green colour to beads of borax and microcosmic salt, before the blowpipe flame.

H_2S gives no precipitate, but *hydro-ammonium sulphide* throws down green hydrated sesquioxide from salts of chrome.

The alkalies throw down green sesquioxide, soluble in excess of alkali: on boiling the solution in alkali the sesquioxide is precipitated again.

They are poisonous, and green, red, yellow, or violet, in colour.

With a salt of lead, a soluble chromate gives a bright yellow precipitate.

When the trioxide or a chromate is boiled with H_2SO_4 and a little alcohol or sugar, they are decomposed, green sesquioxide is formed, which dissolves in the H_2SO_4, but which may be thrown down as hydrated sesquioxide on addition of a little solution of ammonia.

MANGANESIUM.

[*Symbol*, Mn; *Atomic Proportion*, 55; *S. g.* 8·0.]

Occurrence.—This metal is never found native, but, as di- or black oxide (MnO_2), commonly called 'manganese,' is tolerably abundant, and in small quantities is widely diffused through the mineral kingdom.

History.—'Manganese' has been used to colour glass from the earliest times. Scheele in 1774 first recognised it as a peculiar metallic oxide, and Gahn in 1780 first obtained the metal from it.

Preparation and Properties.—By mixing MnO_2 (manganese) with C, and exposing it to the highest furnace white heat. It is a grey-white metal, hard, brittle, with little lustre, and very difficult to melt; oxidises easily in the air, and decomposes cold water slowly, liberating H. The metal is not used in the arts.

Manganesium with Oxygen.

There are five distinct oxides of manganesium; the sixth resembles the magnetic oxide of iron in composition, namely Mn_3O_4 or ($MnO.Mn_2O_3$).

Manganous or Manganesium oxide	$Mn''O''$
Manganic or Manganesium sesquioxide	$Mn'''_2O''_3$
Manganic or Manganesium dioxide	$Mn^{iv}O''_2$
Manganesium trioxide	$Mn^{vi}O''_3$
Manganesium peroxide	$Mn^{vii}_2O''_7$

The most important oxides are the dioxide used to produce chlorine, &c., and the tri- and per-oxides used as disinfectants. The oxide $Mn''O$ is by far the most basylous radical of all; $Mn^{iv}O_2$ is an indifferent oxide, as like all similar dioxides, it does not form any salts; while the tri- and per-oxides are non-basylous bodies, and form salts; namely, the *manganates* and *permanganates*, which are used in the arts and in the laboratory.

MANGANOUS OXIDE OR MANGANESIUM OXIDE.

$[Mn''O'' = 71.]$

MAY be obtained by heating the carbonate ($Mn''CO_3$) to redness in a current of H gas, when it appears as an olive-green body, possessing like the corresponding oxides of Fe'' and Cr'' a great affinity for O, which it absorbs, and becomes converted into a brown oxide, the $Mn''O.Mn'''_2O_3$. It may be precipitated as a white hydrate, by adding an alkali to either the sulphate $Mn''SO_4$, or chloride $Mn''Cl_2$, but it immediately begins to absorb O from the air, and to turn brown. Most of the salts containing Mn'' are of a rose or flesh-colour; they are numerous and may be procured similarly to the corresponding Fe salts.

The sesquioxide ($Mn'''_2O''_3$) is found native in the minerals *braunite* and *manganite*. It is a body of little interest, and the salts containing Mn''' are of no importance.

THE DIOXIDE, DEUTOXIDE, PEROXIDE, BLACK OXIDE, OR MANGANESE.

$[Mn^{iv}O''_2 = 87.]$

THERE are no other salts containing Mn^{iv}.

This oxide is found native in the mineral *pyrolusite.* When mixed with solution of HCl, Cl is liberated (see *C. N.* p. 75); when heated with H_2SO_4, half its O, and when heated red-hot one-third of its O is given off (see *C. N.*

p. 50). HNO_3 has but little action on it. It is largely used in the arts for the preparation of Cl, and in the laboratory as an oxidising agent. It gives an amethyst colour to glass.

The *red brown* oxide, Mn_3O_4, is found in the mineral *hausmannite*. It is produced by heating MnO_2 as in the preparation of O; also it may be made by heating $Mn''CO_3$ in the air.

MANGANESIUM TRIOXIDE

Is not known in a separate state, but is well known as

POTASSIUM MANGANATE.—MINERAL CHAMELEON.

$[K_2MnO_4 = 197.]$

WHEN MnO_2 is heated with KHO in an open dish, the MnO_2 becomes MnO_3, and unites with the KHO; a black mass is the result, which when placed in a small quantity of water partly dissolves, giving a splendid green solution of K_2MnO_4. The salt may be crystallised. It is very unstable; that is, it easily decomposes, for if mixed with any organic matter, as paper, &c., it gives up O, oxidises the organic matter, and deposits a brown oxide of manganesium.

When the green solution of manganate is mixed with a sufficiently *large* quantity of water, it is converted into a bright pink or red solution of potassium permanganate $KMnO_4$.

MANGANESIUM PEROXIDE

Is not known, but the K compound is very important, namely :—

POTASSIUM PERMANGANATE.

$[KMnO_4 = K_2O.Mn_2O_7 \div 2.]$

WHEN the green solution of manganate is greatly diluted with water; or when a solution of manganate is heated, or mixed with a dilute solution of H salt (as HCl), or allowed to stand exposed to the air, a red solution of permanganate is formed, part of the manganate gives up its O to the rest, flakes of brown oxide being precipitated at the same time. The pure salt may be crystallised in long deep red crystals.

The red solution rapidly decomposes organic matter, as paper, &c, brown oxide being precipitated; the colour is also rapidly destroyed by H_2S, SO_2, and other bodies. A solution of this salt is much used in *volumetric* analysis, and both the manganate and permanganate, under the name of Condy's Fluid, are largely used as deodorisers and disinfectants. They are not poisonous.

MANGANESIUM SULPHIDE.

$[Mn''S'' = 87.]$

May be precipitated from a salt containing Mn'' by addition of hydro-ammonium sulphide, as a yellow-red or flesh-coloured precipitate, which rapidly becomes changed to brown if exposed to the air; the presence of traces of iron, cobalt, &c., will cause the precipitate to appear black.

TESTS FOR MANGANESIUM SALTS.

The salts containing Mn'' are the only ones of chemical importance; they are generally of a delicate rose colour.

Before the blowpipe, mixed with sodium carbonate, the salts of manganesium give a very characteristic *green* colour. Heated in a *borax* or microcosmic salt bead, a violet colour is produced by them.

The *alkalies* precipitate white hydrated oxide, which rapidly becomes brown.

The *carbonates* give a stable white precipitate of carbonate, soluble in ammonium chloride.

H_2S gives no precipitate; but *hydro-ammonium sulphide* produces a flesh-coloured precipitate, if the solution of Mn'' is free from Fe'', Co'', &c.

COBALT.

[*Symbol*, Co; *Atomic Proportion*, 59; *S.g.* 8·95.]

OCCURRENCE.—This metal is scarce, and, when found, is always united with either S, arsenic (As), or nickel (Ni), and frequently with all three. It is met with in small quantities in meteorites. The most common ore is *speiss-cobalt* ($CoAs_2$) and *cobalt glance* ($CoS_2.CoAs_2$).

HISTORY.—The roasted ores of cobalt have been used to prepare a blue glass, called *smalt*, since the beginning of the 16th century. The metal was first obtained by Brandt, in 1733.

PREPARATION AND PROPERTIES.—The production of pure oxide of cobalt is a tedious operation. When obtained, the metal may be extracted by heating it to intense whiteness for an hour with C. It is a reddish-grey metal, hard, brittle, and magnetic. It is dissolved by the H salts, forming compounds much used in the arts as pigments (from *pingo*, I paint). The metal is not used in the arts.

COBALT WITH OXYGEN.

COBALT forms two compounds with O, namely—

Cobaltous or Cobalt oxide	$Co''O''$
Cobaltic or Cobalt sesquioxide . . .	$Co'''_2O''_3$

COBALTOUS OR COBALT OXIDE.

[$Co''O'' = 75$.]

THE salts of Co'' are more or less carmine or rose-red when in combination with water, and blue when anhydrous. The oxide is a brown powder, which becomes black if heated in the air. It is decomposed by the H salts, a salt of Co'' being found in solution, from which the hydrated oxide may be precipitated by KHO. With ammonia and cyanogen, this oxide forms a number of complicated and curious double salts.

Zaffre is an impure oxide of Co.

Smalt is a powdered glass, coloured blue by oxide of Co.

Thénard's blue is formed by adding solutions of a mixture of carbonate and phosphate of sodium to a mixture of solutions of cobalt and alum.

Rigman's green is a compound of cobalt oxide with zinc oxide.

COBALTIC OR COBALT SESQUIOXIDE.

$[Co'''_2O''_3 = 166.]$

Is a black powder, soluble in H_2SO_4, with liberation of O, and in HCl, with liberation of Cl, forming salts containing Co''. It is scarcely acted on by HNO_3.

TESTS FOR COBALT SALTS.

BEFORE *the blowpipe* these salts give a bright blue colour to glass, to borax, to microcosmic salt, and to alumina; to magnesia a pink colour, and to zinc a green.

The concentrated solutions of salts of cobalt are blue, and at a certain point of dilution become pink.

H_2S gives no precipitate, but *hydro-ammonium sulphide* a black precipitate of cobalt sulphide.

KHO gives a blue or (in excess) red precipitate of hydrated oxide.

Solution of ammonia throws down a dirty-coloured precipitate, which is soluble in excess, especially if some ammonium chloride be added; when this solution is exposed to the air, it becomes bright pink.

NICKEL.

[*Symbol*, Ni; *Atomic Proportion*, 59; *S.g.* 8·8.]

OCCURRENCE.—Scarce; the most abundant ores are *kupfernickel* ($NiAs$) and *arsenical nickel* ($NiAs_2$), arsenides of nickel. It is also commonly found in the cobalt ores, and is never absent from meteoric iron.

HISTORY.—First obtained in 1751 by Cronstedt.

PREPARATION AND PROPERTIES.—It is very difficult to obtain perfectly pure nickel oxide or nickel oxalate; but from the former the metal is easily obtained by heating it white-hot in a crucible with C, and from the latter by similarly heating it alone. It is a white, ductile metal, difficult to melt, tenacious, and magnetic. It is used in the arts to form *German silver* (an alloy of copper, zinc, and nickel), and is consequently worked in large quantities. Both the metal and its salts in their chemical properties resemble very closely those of cobalt.

NICKEL WITH OXYGEN.

NICKEL forms two compounds with O, namely—

Green Nickelous or Nickel oxide	$Ni''O''$
Black Nickelic or Nickel sesquioxide	$Ni'''_2O''_3$

Solutions of Ni'' are green, from which KHO throws down the hydrated oxide as an apple-green precipitate.

TESTS FOR NICKEL SALTS.

SOLUTIONS of nickel are green.

The alkalies throw down a green hydrate, insoluble in excess.

Ammonia throws down a green hydrate, soluble in excess, from which solution KHO precipitates the hydrate again.

The soluble carbonates give an apple-green *basic* carbonate, of uncertain composition.

H_2S gives no precipitate in solutions of nickel salts, but *hydro-ammonium sulphide* gives a black precipitate of nickel sulphide.

Group B.—Lead Group.

Three metals will be placed in this group, namely—

Lead	Pb
Copper	Cu
Bismuth	Bi

They all oxidise in the air, especially when heated; they do not decompose water at a red heat, neither are they soluble in HCl; their oxides can be reduced by H.

Their solutions may be precipitated by both H_2S and hydro-ammonium sulphide, as black sulphides.

LEAD.

[*Symbol,* Pb (Plumbum); *Atomic Proportion,* 207; *S.g.* 11·40; *Melts at* 330° C.]

Occurrence.—The principal ore of lead is the sulphide called *galena.* Some other salts of lead are found native, mostly crystalline.

History.—Known from the earliest times, and called *Saturn* (♄) by the ancients.

Preparation.—Galena (PbS) is roasted in a reverberatory furnace, with the side-doors open. (See p. 17.) Much of the S burns away as SO_2, while some of it is oxidised to SO_3, which uniting with lead oxide (PbO) formed at the same time, produces $PbSO_4$; part of the galena remains unacted upon, and as large quantities of ore are roasted *at once,* the thorough oxidation of the whole is impossible.

Thus, by the process of roasting, there is formed in the furnace a mixture of PbO, $PbSO_4$, and PbS, in unknown proportion. The roasting being completed, which the workman knows by practice, the side-doors are closed, and the heat increased.

As the heat increases, the doors being closed and the air excluded, the S of the undecomposed PbS takes O from the PbO and $PbSO_4$, forming SO_2 and metallic lead; the former escapes up the chimney, and the latter collects at the bottom of the furnace:—

$$2PbO + PbS = 3Pb + SO_2$$
$$PbS + PbSO_4 = 2Pb + 2SO_2$$

Any ore not decomposed at first is roasted with fresh galena.

Galena generally contains a small quantity of silver, and sometimes gold. To separate these the liquid metal is run into a large iron vessel, and while cooling is briskly stirred. Although pure silver melts at a much higher temperature than lead (see p. 7), *yet an alloy of lead and silver melts at a lower temperature*, consequently, lumps of pure lead solidify first and are removed; the last portions of liquid metal containing all the silver are returned to a reverberatory furnace and roasted with the doors open, when the Pb is converted into PbO, which melts by the heat and is run off, pure Ag remaining at the end of the operation melted at the bottom. The PbO thus formed is again reduced to pure Pb.

PROPERTIES.—A soft, bluish-white metal, easily fusible, with considerable lustre; may be beaten out into very thin leaves, but cannot be drawn to very thin wire, and possesses little tenacity. Volatile. It expands considerably when heated, and does not always contract again to its original form. Tarnishes or oxidises easily in moist air, the oxide when freshly formed being slightly soluble in rain water; and as a solution of lead is poisonous, rain water should never be collected in leaden cisterns; spring water containing salts of lime in solution has little or no action on lead cisterns. Lead is easily soluble in HNO_3, but H_2SO_4 and HCl have little action on it.

LEAD WITH OXYGEN.

LEAD forms two well-known compounds with O, namely—

Plumbic or Lead oxide	PbO
Plumbic or Lead dioxide	PbO_2

The grey scum formed when lead is melted in the air seems to be a suboxide.

The PbO_2, like the corresponding MnO_2, does not form any salts, and *red lead* is a mixture of the two oxides mentioned above.

PLUMBIC OR LEAD OXIDE.—LITHARGE.

$[Pb''O'' = 223.]$

THIS oxide is used in the arts under the name *litharge* (from *λίθος*, a stone, and *ἄργυρος*, silver). It is made during the extraction of silver from lead (see p. 123), and is of a more or less yellow colour according to the mode of preparation. When lead is melted in an iron ladle, the lead becomes covered with a grey coating of oxide, which on continued heating becomes yellow (litharge). A white hydrated lead oxide is precipitated by adding a solution of alkali to a solution of salt of lead; this hydrate is to a slight extent soluble in an excess of alkali, and such a solution forms an excellent test for S. Lead oxide absorbs CO_2 from the air, forming white lead carbonate; it melts at a bright red heat, and easily unites with silica and other bodies, forming fusible mixtures, on which account it is used in glass making and as a glaze for earthenware. (See pp. 65, 98.) It is easily soluble in HNO_3; it acts as a strong basylous radical, and forms an important class of salts.

Red lead ($PbO.PbO_2$) is formed by carefully heating *litharge* in the air, when it absorbs O and becomes red. It is of a brilliant red colour, used as a paint and in the manufacture of glass. It is sometimes called *minium.*

Lead dioxide, binoxide, or puce oxide is made by acting on red lead with HNO_3. This decomposes the PbO contained in it, forming lead nitrate ($Pb2NO_3$), which dissolves, leaving PbO_2.

PLUMBIC OR LEAD NITRATE.

[$Pb2NO_3 = 331.$]

Formed by dissolving either lead or litharge in HNO_3. It is soluble in water, and crystallises in octahedrons, which is commonly opaque or milky looking.

PLUMBIC OR LEAD ACETATES.

Lead oxide (PbO) is dissolved by hydric acetate ($H'.C_2H_3O_2$), a normal salt $Pb''2(C_2H_3O_2)'$ being formed, which from its sweetish taste is commonly called *sugar of lead*, and is very poisonous.

A solution of sugar of lead can dissolve lead oxide, PbO, a curious tribasic salt, called in medicine *Goulard's water*, being formed ($Pb''2(C_2H_3O_2)'.2PbO$). This salt most easily absorbs CO_2 from the air, becoming milky from formation of $PbCO_3$, the normal salt $Pb2(C_2H_3O_2)$ being left in solution.

PLUMBIC OR LEAD CARBONATE.

[$PbCO_3 = 257.$]

Is frequently found native, in long, white, needle-shaped crystals. May be formed by precipitating a solution of sugar of lead by sodium carbonate. It is a white, heavy powder, insoluble in water.

White lead is an impure lead carbonate used as a paint. Lead carbonate precipitated by sodium carbonate from sugar of lead, does not answer as a paint, because it is slightly crystalline, and does not *cover* sufficiently. In commerce the white lead is made by a very slow process, and is perfectly amorphous. This latter preparation covers the work with a smooth coating of paint. Small pots containing a little vinegar at the bottom, have a thin piece of lead made to rest inside them just over the vinegar. Each pot is then covered, and numbers of them arranged in rows; they are now covered with old tan (refuse oak bark), and more pots arranged in rows on the others; these are again covered by tan, and the process continued. Vinegar contains hydric acetate ($H.C_2H_3O_2$), the loose tan allows air to circulate, and the tan by decomposing produces CO_2, The lead in each pot oxidises, and the oxide is dissolved by the hydric acetate, which is volatile, and so reaches the lead, although it does not touch it. The lead acetate $Pb\,2(C_2H_3O_2)$ thus formed, dissolves lead oxide as this latter becomes formed, and so the salt $Pb\,2(C_2H_3O_2).2PbO$ is produced (see *Lead acetates*); the last salt easily unites with $2CO_2$ from the tan, and forms $2PbCO_3$. The normal lead acetate left, then dissolves more PbO, and so the process gradually continues for weeks. Finally, on the tan being removed, and the pots opened, the $PbCO_3$ is taken out, washed, ground to powder, and sold as white lead.

White lead is often adulterated with barium sulphate and other white powders. Painters and workmen who handle lead salts are subject to a peculiar disease called *lead* or *painter's colic.*

PLUMBIC OR LEAD SULPHATE.

[$PbSO_4$.]

Is a white heavy *insoluble* powder. It may be distinguished from barium sulphate by giving a black colour *with* H_2S.

PLUMBIC OR LEAD SULPHIDE.

[$PbS = 223$.]

This salt occurs as a native ore of lead in *galena*; it is frequently found crystallised in dark blue-black cubes; when freshly broken the surface possesses considerable lustre. (See pp. 1, 13.) It may be precipitated as a black hydrate by adding H_2S or hydro-ammonium sulphide to any salt of lead. (See *C. N.* p. 88.) The hydrate is decomposed by HNO_3 into insoluble lead sulphate ($PbSO_4$), it is also decomposed by a hot and concentrated solution of HCl.

Lead sulphide is decomposed when heated with metallic iron.

PLUMBIC OR LEAD CHLORIDE.

[$PbCl_2$.]

Is chiefly remarkable as being one of the only three chlorides which are insoluble in water. (See p. 36.) It may be precipitated by adding HCl or NaCl to a solution of a lead salt.

PLUMBIC OR LEAD IODIDE.

[PbI_2.]

Is a beautiful yellow salt with difficulty soluble in water. May be formed by adding a solution of KI to one of lead acetate.

TESTS FOR LEAD SALTS.

The salts of lead are poisonous.

Before the blowpipe on charcoal, salts of lead are reduced to a malleable globule of metal, which is surrounded on the charcoal by a ring of yellow oxide.

The alkalies precipitate from solutions of lead salts white hydrated oxide.

The soluble carbonates precipitate white lead carbonate.

Dihydric sulphate precipitates white lead sulphate.

Hydric chloride precipitates white lead chloride, which is *not* changed by solution of ammonia. This chloride is slightly soluble in water.

Potassium iodide gives a bright yellow precipitate of lead iodide.

Potassium chromate gives a yellow precipitate of lead chromate.

H_2S and *hydro-ammonium sulphide* give a black precipitate of lead sulphide insoluble in excess of the precipitant or of alkalies. With this exception all the insoluble salts of lead are more or less soluble in KHO.

Lead is easily reduced from its salts by a piece of zinc (the *lead tree*), the zinc becoming dissolved in place of the lead.

COPPER.

[*Symbol*, Cu (Cuprum); *Atomic Proportion*, 63·5; *S.g.* 8·95.]

OCCURRENCE.—Frequent; sometimes native, but the commonest ore of copper is *copper pyrites* or double sulphide of copper and iron ($CuFeS_2$). In Australia the principal ore of copper is *malachite*, an impure carbonate.

HISTORY.—Known from the earliest times, and called Venus (♀) by the ancients. (See p. 32.)

PREPARATION.—The process by which Cu is obtained from copper pyrites ($CuFeS_2$) appears a most complicated one to a beginner. It consists of at least five distinct processes:—

1st. *The ore is roasted* either in the air or in reverberatory furnaces. By this process much of the S is got rid of as SO_2, and as Cu has a stronger affinity for S than Fe has, the greater part of the Fe becomes converted into Fe_2O_3, the Cu remaining in combination with S, as Cu_2S.

2nd. *The roasted ore is smelted in a furnace* with a slag containing much silica. Fluor spar is often added. The Fe_2O_3 and the earthy matters of the ore are by this means to a great extent got rid of; a dicopper sulphide (Cu_2S) containing a little iron sulphide, sinks through the slag to the bottom of the furnace, from whence it is run out, and is called *coarse metal.*

3rd. *The impure copper sulphide* (*coarse metal*) obtained from the furnace as above, *is again roasted for* 24 *hours* in a reverberatory furnace to further oxidise the iron sulphide which it still contains.

4th. *The re-roasted ore is again smelted with slag* as in process 2, to remove the Fe_2O_3 formed by process 3; the copper sulphide (Cu_2S), which runs from the furnace after this process is nearly pure, and is called *fine metal.*

5th. The copper sulphide obtained by process 4 is again roasted in the reverberatory furnace for some hours, at such a heat that part of the dicopper sulphide (Cu_2S) becomes converted into copper oxide (CuO), a mixture of $Cu_2S + CuO$ being obtained. The heat is now increased, when the mass melts, and the sulphide and oxide decompose one another thus:—

$$Cu_2S + 2CuO = SO_2 + 4Cu$$

The metallic Cu sinks to the bottom and is run out of the furnace (blistered copper). To obtain the best copper the metal thus procured requires to be further refined, in which process any oxide is finally removed by stirring the melted metal with a wooden pole. This operation is called *poling.*

The Cu from *malachite* can be reduced by C alone, assisted by a little slag to remove impurities.

The water in the neighbourhood of many copper mines contains Cu, which it has dissolved as copper sulphate ($CuSO_4$) obtained from the gradual oxidation of pyrites. The Cu may be precipitated from such water by placing scraps of clean iron in it, thus: $CuSO_4 + Fe = FeSO_4 + Cu$. Near copper works, as at Swansea, a white deleterious smoke (copper smoke) is constantly being given off which destroys all vegetation near, and contains arsenic obtained from the ore.

PROPERTIES.—A metal with a well-known red colour (copper colour). It is rather hard, but very tenacious, ductile, and malleable, possessing great lustre. It is one of the best conductors of heat and electricity. If heated to nearly its melting point it is brittle, and if heated in the air becomes covered with black scales of oxide (CuO). In moist air often becomes covered with green carbonate. Does not decompose water, and is not soluble in cold H_2SO_4 or boiling HCl, but is easily dissolved by HNO_3.

It is used for money, coating ships, making kettles and other utensils. It forms part of the alloys called brass, bronze, bell-metal, German silver, &c., and its salts are used as paints and in medicine. They are very poisonous.

COPPER WITH OXYGEN.

COPPER forms two well-known compounds with O, namely—

Red Cuprous or Dicopper oxide . . Cu'_2O''.
Black Cupric or Copper oxide . . $Cu''O''$.

The CuO is the oxide of chief importance; its salts are generally green or blue when hydrated, and white when anhydrous.

CUPROUS OR DICOPPER OXIDE.

[$Cu'_2O'' = 143$.]

THIS oxide occurs native, crystallised in octahedrons, and called *octahedral copper ore*; it is of a red or yellow colour. It can be prepared artificially by adding solution of

KHO to a mixture of copper sulphate and grape sugar, and gently warming the blue solution, when a copious pale yellow or red precipitate of Cu_2O is formed.

CUPRIC OR COPPER OXIDE.

[$Cu''O'' = 79{\cdot}5.$]

A BLACK powder, forms the black scales when Cu is heated in the air. Is generally obtained for use in the laboratory by heating copper nitrate ($Cu2NO_3$) in a crucible.

It may be precipitated as a beautiful blue hydrate ($CuO.H_2O$) by adding KHO to a solution of any of its salts. This hydrate loses its water and becomes black when boiled. The hydrate is soluble in ammonia forming a beautiful deep purple blue coloured liquid.

The oxide is hygroscopic; it is easily reduced to the metallic state when heated, by either H or C. (See *C. N.* p. 54.)

CUPRIC OR COPPER SULPHATE.—BLUE VITRIOL.

[$CuSO_4.5H_2O = 159{\cdot}5 + 90.$]

MAY be formed by boiling Cu in dihydric sulphate (see *C. N.* p. 83); on the large scale by roasting copper sulphide (see p. 102) and lixiviating the mass in water. Forms large clear blue crystals soluble in 4 parts of cold water. When heated it loses its water of crystallisation, becoming white, and has then so great an affinity for water that it is frequently used to deprive other bodies of it. It is largely employed in calico printing.

CUPRIC OR COPPER SULPHIDE.

[$CuS = 95{\cdot}5.$]

MAY be formed by heating Cu in vapour of S (see p. 11), it is also thrown down of a dark brown or black colour whenever H_2S or hydro-ammonium sulphide is added to a solution of a salt of Cu''. There is also a sulphide Cu'_2S formed during the process of copper smelting.

CUPRIC OR COPPER NITRATE.

$[Cu2NO_3+6H_2O=187{\cdot}5+108.]$

This salt is easily obtained by dissolving Cu in hydric nitrate (see *C. N.* p. 69). The salt is deliquescent, and is used for fireworks. It is of a fine blue colour.

CARBONATES OF COPPER.

There are several impure hydrated carbonates of copper found native, of which *malachite* is the most famous; they all contain hydrated oxide. The pure carbonate has never been formed. When any salt of Cu'' is precipitated by sodium carbonate a green or blue precipitate of a mixture of hydrated oxide and carbonate of copper is obtained.

Blue copper ore	$2(CuCO_3)+CuO.H_2O$
Green copper ore or *malachite* .	$(CuCO_3)+CuO.H_2O$

There are two *chlorides of copper* corresponding to the two oxides. The one, *white* $Cu'_2Cl'_2$, formed when Cu is heated in Cl gas: it is *insoluble* in water but soluble in solution of HCl. The other, *green* and may be formed by decomposing the hydrated oxide (CuH_2O_2) with HCl, when $Cu''Cl_2$ is formed; it is *soluble* in water.

TESTS FOR COPPER SALTS.

The red oxide Cu'_2O'' gives a beautiful blood red colour to glass.

The black oxide $Cu''O''$ gives a beautiful green colour to glass; when heated on charcoal before the blowpipe and covered with sodium carbonate to act as a flux, a globule of metallic copper may be obtained.

The salts of copper are very poisonous.

The *alkalies* KHO and NaHO precipitate pale blue hydrated oxide, which turns black when boiled.

Ammonia precipitates pale blue or green hydrated oxide, which dissolves in excess of ammonia to a beautiful deep purple blue colour.

The *soluble carbonates* precipitate green hydrated oxy-carbonates, which are soluble in excess of ammonium carbonate.

Potassium ferrocyanide gives a claret-coloured precipitate in solutions of copper salts, insoluble in HCl.

A *strip of clean iron* placed in solution of copper, to which a few drops of a solution of H salt has been added, precipitates bright metallic Cu, the iron being dissolved in place of the copper.

H_2S and *hydro-ammonium sulphide* precipitate black copper sulphide.

BISMUTH.

[*Symbol*, Bi; *Atomic Proportion*, 210; *S.g.* 9·8.]

OCCURRENCE, HISTORY, AND PREPARATION.—Bismuth is only found in a few places, and is not very abundant; the chief supply comes from Saxony. Was discovered in 1529 by Agricola. Is mostly found native, and simply requires to be melted out of the matrix in which it occurs; is also occasionally found as sulphide. It is used as a medicine, and as a paint for the face.

PROPERTIES.—A reddish-white metal, hard, brittle, volatile, and crystalline; its crystals are really rhombs although they look like cubes. It forms alloys, the chief of which is *fusible metal*, composed of 2 parts of Bi, 1 of Pb, and 1 of Sn. The metal easily oxidises when heated in the air, burning with a bluish flame. It is also rapidly dissolved by hydric nitrate, forming $Bi'''3(NO_3)' + 5H_2O$.

Bismuth with Oxygen.

Two compounds of Bi with O are well known, others being doubtful.

Bismuthic or Bismuth sesquioxide . .	$Bi'''_2O''_3$
Bismuthic or Bismuth peroxide . .	$Bi^v{}_2O''_5$

The sesquioxide Bi_2O_3 may be formed by (1) heating the metal in the air, or (2) by heating the basic nitrate to low redness. It is a yellowish powder, insoluble in water, easily melts, and may be reduced by H or C to metallic Bi. A white hydrated oxide may be precipitated from a solution of bismuth salts by ammonia.

The brown Bi_2O_5 is of little importance. It may be obtained by passing Cl through Bi_2O_3 suspended in solution of KHO.

BISMUTHIC or BISMUTH NITRATE.

$[Bi'''3(NO_3)'+5H_2O = 396+90.]$

This is a soluble salt of Bi, of much importance. It is easily obtained by dissolving Bi in HNO_3, from which solution the salt may be obtained in fine large crystals.

There is a peculiarity with the salts of Bi—namely, *if much water be added to them, they turn milky, a white, basic, insoluble salt falling as a precipitate.* Thus, when solution of $Bi3(NO_3)$ is thrown into water, a white basic oxynitrate, called *tris-nitrate of bismuth*, is thrown down $(Bi'''[(NO_3)'O''].H_2O)$, a small quantity of bismuth remaining in solution. It is used in medicine.

BISMUTHIC or BISMUTH CHLORIDE.

$[Bi'''Cl'_3 = 316{\cdot}5.]$

Is formed (1) when the metal is dissolved in a mixture *of HCl* and HNO_3, or (2) when Bi_2O_3 is dissolved in HCl. *When thrown into water*, like the nitrate, it decomposes into an oxychloride salt, which is insoluble in water, and much used as a face powder (pearl white)—

$$3(Bi'''Cl'_3) + \text{aqua} = [Bi'''_3(Cl'_3O''_3).H_2O] + 6HCl$$

The chloride of bismuth resembles the chloride of antimony, in forming a white precipitate when thrown into water; this reaction is very characteristic of the salts of the two metals. The white precipitates may easily be distinguished from one another, for the antimony salt is soluble in hydric tartrate, while the bismuth salt is not.

BISMUTHIC OR BISMUTH SULPHIDE.

$[Bi'''_2S''_3 = 516.]$

OCCURS native. When solutions of Bi are treated with H_2S, a black precipitate of sulphide is formed.

TESTS FOR BISMUTH SALTS.

Before the blowpipe on charcoal, they are reduced to a brittle globule of metallic Bi.

The alkalies precipitate white hydrated oxide.

When thrown into a large quantity of water, a white insoluble oxysalt is produced. This precipitate is sometimes difficult to produce in presence of large quantities of free H salts. The oxysalts are not soluble in dihydric tartrate ($H_2.C_4H_4O_6$).

Potassium chromate gives a yellow precipitate of bismuth chromate, insoluble in KHO, but soluble in dilute solution of HNO_3.

H_2S throws down a black sulphide insoluble in *hydro-ammonium sulphide.*

GROUP C.—TIN GROUP.

THREE metals will be considered in this group, namely—

Tin	Antimony	Arsenic or Arsenicum

The following metals also belong to this group to the extent, that, like those above, they form distinct *non*-basylous oxides:—

Titanium
Molybdenum
Tungsten
Vanadium
Tantalum
Columbium

All these metals form a class of salts called, respectively, *stannates, antimoniates, arseniates, titanates, molybdates, tungstates, vanadiates.* The metals tantalum and columbium also form similar salts; little, however, is known about them.

Bismuth forms a non-basylous oxide, on which account it might be classed with the Tin group; its salts also bear some relation to those of antimony; in other respects, however, it does not belong to the Tin group.

The metals of this group, with the exception of arsenic, decompose water at a red heat. They do not liberate H when in contact with *non*-basylous hydric sulphate (H_2SO_4), or hydric chloride (HCl) when cold. They liberate H in contact with *basylous* hydro-potassium oxide (KHO). They each form more than one compound with S, the *per* or higher sulphide *being soluble in hydro-ammonium sulphide*—that is, their persulphides form double and soluble salts with the sulphides of the alkalies, many of which may be crystallised similarly to the double O salts. H_2S, *in presence of* HCl, *precipitates these metals from their solutions.*

TIN.

[*Symbol,* Sn (Stannum); *Atomic Proportion,* 118; *S.g.* 7·30; *Melting Point,* 228° C.]

OCCURRENCE.—In few localities; almost entirely as *Tinstone* ($Sn^{iv}O''_2$). It is found chiefly in Cornwall, which has been noted from very early times for its tin ores; it is also found in Malacca. It does not occur native.

History, &c.—Called by the ancients Jupiter (♃). Pliny says the Phœnicians traded with Britain for tin, which they used in the manufacture of bronze. It is prepared easily by first crushing, roasting, and washing the ore (see p. 15), and then heating it in a reverberatory furnace with coal, and a little lime to act as a flux. It is afterwards remelted very carefully, and the purer portions, which melt first, separated from an impure alloy containing As, Pb, and Cu, &c.

Tin is a silver-white, very lustrous metal; does not tarnish or oxidise in the air; is soft, very malleable (tin-foil), and when in bars crackles if bent. It is not volatile, and burns in the air to oxide, SnO_2.

Tin is not dissolved by cold but is by boiling HCl, with liberation of H, and HNO_3 converts it into a white insoluble powder, the SnO_2. The salt HNO_3 has a similar action on antimony. (See p. 141.)

Tin forms part of the alloys called bronze, bell-metal, pewter, britannia metal, and solder. It is also used to coat clean sheet iron, which is then called *tin-plate.*

Tin with Oxygen.

There are two principal and well-known oxides of tin, namely—

Stannous or Tin oxide . . .	$Sn''O''$
Stannic or Tin dioxide . . .	$Sn^{iv}O''_2$

The $Sn''O''$ is a violet-black powder of little importance; a white hydrated oxide may be precipitated. The salts of Sn'', especially the chloride $Sn''Cl_2$ are however of importance.

STANNIC or TIN DIOXIDE.

$[SnO_2 = 150.]$

This is the most important compound of tin.

1. It occurs native as *tinstone.*
2. It is formed when Sn is burned in the air.

3. It is formed when Sn is acted on by (HNO_3).

4. By precipitating $SnCl_4$ with solution of ammonia or a carbonate.

Properties.—The native mineral is brown and very hard. It is insoluble in H salts, but soluble after fusion with KHO as potassium stannate. Prepared by process 2 or 3 it is a white powder *insoluble* in H salts, but soluble in the alkalies; it forms the *putty powder* used to polish plate, and to give white opacity to enamels. As prepared by process 4 it is *soluble* in H salts, but if heated it is converted into the insoluble oxide. The hydrated oxide is a non-basylous radical forming a class of salts called *stannates.*

The salts made from the soluble SnO_2 have a formula corresponding to H_2SnO_3, and those from the insoluble variety $H_2Sn_5O_{11}$. *The stannates* are much used by the *dyers.* The *tin-prepare liquor* is a sodium stannate.

TIN WITH SULPHUR.

There are two well-known compounds of Sn with S, corresponding to the two oxides, namely—

Brown Stannous or Tin sulphide . . .	$Sn''S''$
Yellow Stannic or Tin disulphide . . .	$Sn^{iv}S''_2$

The brown sulphide is produced by passing H_2S through a solution of Sn''; the precipitate is soluble in hydro-ammonium sulphide, and a strong solution of HCl decomposes it, forming $Sn''Cl_2$ and liberating H_2S.

STANNIC OR TIN DISULPHIDE.

$[Sn^{iv}S''_2 = 182.]$

This sulphide, called *mosaic gold,* from its beautiful yellow colour, is very like the sulphide of As in appearance. It may be precipitated from a solution of Sn^{iv} by H_2S. It is, like sulphide of As, soluble in hydro-ammonium sulphide. When the two sulphides are mixed together they are difficult to separate perfectly.

TIN WITH CHLORINE.

THERE are two chlorides of tin corresponding to the two oxides, namely—

Stannous or Tin chloride . . .	$Sn''Cl'_2$
Stannic or Tin tetrachloride . .	$Sn^{iv}Cl'_4$

STANNOUS OR TIN CHLORIDE.

$[Sn''Cl'_2 = 189.]$

Is commonly prepared by boiling Sn in solution of HCl, when H is liberated, $SnCl_2$ being formed and dissolved; it may be crystallised ($SnCl_2.2H_2O$) when it forms the *salts of tin* of the dyer. By heat the $2H_2O$ may be removed. When these crystals are dissolved in a large quantity of water they decompose, forming an oxychloride and making the *solution milky, much in the same way as the antimony and bismuth salts do.* This chloride, containing Sn'', has a great power or desire to attract O or Cl and pass into Sn^{iv}, on which account it is frequently used to reduce other bodies; thus, salts of Hg are reduced to the state of metal by it. Gold chloride is also decomposed by it, and a purple precipitate of gold and tin, called *purple of Cassius*, formed.

STANNIC OR TIN TETRACHLORIDE.

$[Sn^{iv}Cl'_4 = 260.]$

THIS chloride is a thin, volatile, colourless liquid fuming in the air, and formerly called *spirit of Libavius.* (See p. 18.) It is obtained by distilling Sn with four times its weight of mercury dichloride ($HgCl_2$) :—

$$2HgCl_2 + Sn = 2Hg + SnCl_4$$

Care must be taken to have the material free from moisture, as water decomposes it.

It may also be made by passing Cl over melted Sn.

This chloride may be dissolved in a small quantity of water. An impure solution of tetrachloride may also be made by dissolving Sn in a mixture of HCl and HNO_3, or in a mixture of HNO_3 and ammonium chloride.

TESTS FOR TIN SALTS.

METALLIC zinc precipitates Sn in the metallic form from all its solutions.

Before the blowpipe, on charcoal, with a little $NaCO_3$ for a flux, the salts are reduced, a globule of metallic Sn being formed.

Salts containing Sn''.

The alkalies KHO and NaHO precipitate white hydrated oxide, soluble in excess. The solution is partly decomposed by boiling. This precipitate is not soluble in ammonia.

H_2S precipitates brown sulphide of tin, *soluble* in hydro-ammonium sulphide.

With a dilute solution of gold chloride a purple precipitate is formed (purple of Cassius).

Salts containing Sn^{iv}.

The *alkalies* precipitate white hydrated oxide, soluble in excess.

H_2S precipitates yellow sulphide of tin, *soluble* in hydro-ammonium sulphide.

ANTIMONY.

[*Symbol,* Sb (Stibium); *Atomic Proportion,* 122; *S.g.* 6·70; *Melts at* 450° C.]

OCCURRENCE.—Tolerably abundant; generally as grey or black sulphide (*grey antimony ore*), sometimes native.

HISTORY.—Discovered by a monk named Basil Valentine. He gave a salt of antimony to some pigs, and found that they got fat on it; he then gave some to his brother monks, to make them fat, but the monks became very sick and ill, in consequence of which he called it antimony (antimoine) (from *ἀντί*, against, and *μοναχός*, a monk). The salt called *butter of antimony* is now sometimes given horses to make them look sleek, but its use requires care, as all the salts of antimony are poisonous. Many of them however are used in medicine.

Pure antimony was first obtained in the fifteenth century.

Preparation.—The native sulphide is first melted out of the matrix in which it occurs, and is collected.

1. The sulphide ($Sb'''_2S''_3$) is melted with scraps of iron:

$$Sb_2S_3 + Fe_3 = 3FeS + Sb_2.$$

2. A mixture of sulphide, crude tartar, and nitre is thrown into a red-hot crucible.

3. The sulphide is roasted in the air until the S is burnt away as SO_2 and the antimony converted into oxide. The so-formed impure oxide is smelted with charcoal and a little sodium or potassium carbonate is added to act as a flux. The metal collects under a layer of slag.

Properties.—A tin-white, lustrous, crystalline and very brittle metal, so brittle that it may be powdered in an earthen mortar; does not easily tarnish in the air, melts readily, but is not so volatile as As; when heated, burns in the air, forming a white oxide ($Sb'''_2O''_3$).

The salts HCl and H_2SO_4 do not dissolve Sb *when cold*, but do when heated.

HNO_3 attacks the metal, converting it (like Sn) into a white insoluble oxide $Sb^v_2O''_5$.

The alloys of antimony are important—*type metal* composed of Pb and Sb, with sometimes Sn, *pewter*, *Britannia metal*, and the old *pocula emetica*. The Sb causes the alloy to expand slightly in cooling, thus giving great sharpness to the casting.

ANTIMONY WITH OXYGEN.

THERE are two well-known compounds of Sb with O, namely—

Antimonious oxide or Antimony sesquioxide . . .	$Sb'''_2O''_3$
Antimonic oxide or Antimony peroxide	$Sb^v_2O''_5$.

There is a grey body, probably a mixture of oxides.

ANTIMONIOUS OXIDE OR ANTIMONY SESQUIOXIDE.

[$Sb'''_2O''_3 = 292$.]

PREPARATION.—This oxide may be formed—

1. By burning Sb in the air.

2. By decomposing Sb_2S_3 by an only moderately strong solution of HNO_3.

3. By pouring antimony trichloride ($SbCl_3$) into water, with addition of a little $NaCO_3$ to complete the precipitation; when a milky precipitate of hydrated oxide is formed.

PROPERTIES.—Either as a white powder (hydrated oxide) or clear lustrous crystals of anhydrous oxide obtained by sublimation. *Insoluble* in water and HNO_3, *soluble* in strong solution of HCl and KHO, also in potassium ditartrate, with formation of *tartar emetic.*

It is dimorphous, crystallising in octahedrons and rhombic prisms. It is easily reduced to Sb by C or H. It acts the part of a weak *non*-basylous radical.

ANTIMONIC OXIDE OR ANTIMONY PEROXIDE.

[$Sb^v_2O''_5 = 324$.]

PREPARATION.—This oxide may be produced—

1. By acting on Sb by very strong solution of HNO_3.

2. By pouring antimony pentachloride ($Sb^vCl'_5$) into water.

3. By melting Sb with nitre (KNO_3), when potassium *antimoniate* is the result.

PROPERTIES.—The hydrated oxide is a white powder, the anhydrous oxide pale yellow. It loses O when strongly heated, forming a grey oxide.

It is insoluble in H salts, but is soluble in the alkalies, forming a class of salts called *antimoniates*, the *antimoniate of sodium* being an insoluble salt, the only insoluble salt of sodium known, that can be used as a test for that metal.

ANTIMONY WITH SULPHUR.

Two sulphides of Sb are known, corresponding to the two oxides, namely—

Antimonious sulphide or Antimony sesquisulphide . $Sb'''_2S''_3$
Antimonic sulphide or Antimony persulphide . $Sb^v_2\,S''_5$

The latter is called *gold sulphur*, and forms a class of double sulphur salts; the double sulphide of sodium and antimony being $Na'_3Sb^v.S''_4+9H_2O$.

ANTIMONIOUS SULPHIDE OR ANTIMONY SESQUI-SULPHIDE.

$[Sb'''_2S''_3 = 340.]$

THIS body is found as the grey-black native sulphide. Precipitated by H_2S from a salt containing Sb''' as from the trichloride ($Sb'''Cl'_3$), it is of a deep orange-red colour, and if made to crystallise turns to grey-black.

ANTIMONY WITH CHLORINE.

THERE are two chlorides of antimony, corresponding to the two oxides and the two sulphides, namely—

Antimonious chloride or Antimony trichloride . $Sb'''Cl'_3$
Antimonic chloride or Antimony pentachloride . $Sb^v\,Cl'_5$

The trichloride ($SbCl_3$) is easily formed by dissolving the native Sb_2S_3 in strong solution of HCl; large quantities of H_2S are given off (see *C. N.* p. 87):—

$$Sb'''_2S''_3 + 6HCl = 2(Sb'''Cl'_3) + 3H_2S$$

Like the chlorides of Bi and Sn, it is decomposed when thrown into water, a curious white oxychloride being precipitated, called *powder of algaroth*, $2SbCl_3.5Sb_2O_3$. If this powder be dried and heated in a retort, the chloride $SbCl_3$ distils over as a thick liquid, which sets when cold to a body like common butter, called *butter of antimony*; the oxide Sb_2O_3 remains in the retort. (See p. 18.)

The trichloride is a volatile, fusible, deliquescent, crystalline solid, soluble in solution of HCl or in a *small* quantity of water.

The pentachloride ($SbCl_5$) may be prepared by—

1. Passing dry Cl over Sb gently heated. The metal takes fire, and the chloride distils over.

2. By passing Cl through the trichloride.

PROPERTIES.—A colourless, heavy, volatile, smoking liquid. Decomposes when thrown into a *large* quantity of water, forming HCl and $Sb^{v}_2O''_5$.

ANTIMONIURETTED HYDROGEN.

[H_3Sb.]

THIS gas has never been obtained pure. It is given off, together with free H, whenever Zn, a salt of antimony, and H_2SO_4 are mixed together.

If this gas is passed through a glass tube heated at one place, the gas is decomposed at that point, and the Sb deposits on the glass as a bright black ring.

When the gas is burned in the air, the flame gives white fumes of oxide; and if a cold body—say a clean white pot—be held in the flame, Sb is deposited as a black spot where the flame touches. This is called *Marsh's test.*

TESTS FOR ANTIMONY SALTS.

BEFORE the blowpipe, the salts of antimony are either partly or entirely volatile.

They are poisonous.

Except the potassio-tartrate (tartar emetic), most of the salts of antimony are decomposed, or *turn milky* when largely diluted with water. This precipitate or milkiness is dissolved by either hydric tartrate or potassium bitartrate.

The alkalies precipitate (in absence of hydric tartrate) a white precipitate, soluble in excess.

Ammonia and the *soluble carbonates* give a white precipitate, nearly insoluble in excess.

H_2S gives a characteristic *orange-red* precipitate with salts of antimony, which is soluble in *hydro-ammonium sulphide.*

MARSH'S TEST.—By antimoniuretted hydrogen. The black ring or spots are easily soluble in hydro-ammonium sulphide, and are *insoluble in solution of chloride of lime.*

REINSCH'S TEST.—By boiling strips of clean copper in a solution of antimony to which a few drops of solution of HCl have been added. The copper becomes covered with a dark coating of metallic antimony.

ARSENIC.

[*Symbol*, As; *Atomic Proportion*, 75; *S.g.* 5·8.]

OCCURRENCE.—Widely distributed, in small quantities, through many metallic ores; often found native, but mostly in combination with Fe, Co, Ni, and S.

HISTORY.—Several compounds of this metal appear to have been known and used as poisons from the earliest historic times. Brandt, in 1733, first carefully studied it. It was formerly known in commerce under the name of *cobaltum.* Its salts are used in medicine.

PREPARATION.—There are several ways by which As is obtained:—

1. By subliming any ore which contains the metal in its native state.

2. By subliming *mispickel* (FeAsS), when As goes off, FeS remaining.

3. By roasting any ore of arsenic, when the As sublimes as As_2O_3, which is carefully condensed in large rooms. The As_2O_3 is then reduced by C and heat.

PROPERTIES.—A brilliant, dark steel-grey, lustrous metal, which soon tarnishes in the air; it is nearly as brittle as Sb, and *far more volatile.* When volatilised in the air (as on a piece of charcoal before the blowpipe), it emits a peculiar, garlic-like *smell*; as *all vapours* of arsenic are *highly poisonous,* care must be exercised in experimenting with it.

This metal resembles P very closely, forming classes of salts, the *arsenates,* which are very like the phosphates in composition.

It is easily dissolved by HNO_3, with formation of $As^{v}{}_2O''{}_5$; it is soluble in a solution of bleaching powder, and is scarcely acted on by HCl.

ARSENICUM WITH OXYGEN.

ARSENICUM forms two compounds with O, corresponding to the oxides of N, P, Sb, and Bi, namely:—

Arsenious oxide or Arsenic sesquioxide . .	$As'''{}_2O''{}_3$
Arsenic oxide or Arsenic peroxide . . .	$As^{v}{}_2O''{}_5$

When As is exposed to the air, a black film forms on it, which may be an oxide ($As'{}_2O''$); it is partly soluble in water, forming the *fly water* (poison) of the apothecary's shop.

The oxide As_2O_5 resembles very closely the body P_2O_5. They are both *tri*valent non-basylous radicals, and unite with 3 molecules of water, forming the normal H salts; thus:—

$$As^{v}{}_2O''{}_5 + 3H_2O = 2(H_3.AsO_4), \text{ the trihydric arsenate}$$
$$P^{v}{}_2O''{}_5 + 3H_2O = 2(H_3.PO_4), \text{ the trihydric phosphate}$$

Again, the analogy between the two non-basylous oxides P_2O_3 and As_2O_3 may be thus shown:—

$As'''_2O''_3 + 3H_2O = 2(H_3.AsO_3)$, trihydric arsenite
$P'''_2O''_3 + 3H_2O = 2(H_3.PO_3)$, trihydric phosphite

The fact that As_2O_3 does not crystallise with 3 molecules of H_2O, and that the three H in H_3PO_3 have not as yet been substituted by metal, does not alter the position and analogy these salts bear to one another. (See *C. N.* p. 91.)

Further, N occurs as $N^{v}_2O''_5$ and as $N'''_2O''_3$; these oxides have *never yet* been obtained in combination with more than 1 molecule of H_2O; thus:—

$N^{v}_2O''_5 + H_2O = 2(H.NO_3)'$, mono-hydric nitrate
$N'''_2O''_3 + H_2O = 2(H.NO_2)'$, mono-hydric nitrite

These, therefore, correspond to the *mono*-hydric phosphate $(HPO_3)'$. (See p. 64; also see *C. N.* p. 92.)

From analogy, however, it is not impossible but that some day a body $(H_3NO_4)'''$ will be discovered corresponding to $(H_3PO_4)'''$, as well as the present $(HNO_3)'$, which corresponds to $(HPO_3)'$.

ARSENIOUS OXIDE OR ARSENIC SESQUIOXIDE.

$[As'''_2O''_3 = 198.]$

THIS body is one of the oldest poisons known. It is also the commonest form in which As is met with, and is found in the market under the names of *white arsenic*, *glass of arsenic*, and often simply called *arsenic*.

PREPARATION.—The metal As easily unites with O:—

1. By roasting arsenical pyrites (FeAsS) in a current of air, and collecting the vapours of As_2O_3 which escape.
2. By roasting the native metallic As in a current of air.
3. By dissolving metallic As in a dilute solution of HNO_3.

PROPERTIES.—It occurs in two forms, namely:—

1. When first sublimed in large quantities, as a clear, vitreous, and transparent solid, like glass. This, however,

by keeping slowly changes to an opaque mass, like white porcelain, its composition remaining the same.

2. Sublimed carefully in very small quantities, or crystallised from a solution of HCl, it is deposited in beautiful lustrous transparent regular octahedra.

It is only slightly soluble in water, more so in HCl, and easily taken up by alkalies, an *arsenite* of the alkali being formed. These arsenites are *all soluble in water*, do not form crystals, and are very poisonous. *Fowler's solution*, used in medicine, is an arsenite of potassium. The arsenites have a formula corresponding to $H'_3.(AsO_3)'''$.

The arsenites of the alkaline earths are *insoluble* in water, on which account *lime* and *magnesia* are very useful *antidotes* in case of poisoning by arsenic (from *ἀντί*, against, and *δίδωμι*, I give).

Freshly precipitated hydrated oxide of iron is also given as an antidote. An insoluble *arsenate* of iron is formed (Fe_32AsO_4).

This non-basylous oxide also forms salts with the heavy basylous oxides, which are chiefly insoluble in water, but poisonous; thus:—

Scheele's Green, or arsenite of copper ($CuH.AsO_3$), formed by adding copper sulphate to a hot solution of 1 part of As_2O_3 and 3 of KCO_3. *Arsenite of silver* ($Ag_3.AsO_3$), a *canary-yellow* coloured precipitate.

The vapour of As_2O_3 has no smell.

ARSENIC OXIDE OR ARSENIC PEROXIDE.

[$As_2O_5 = 230.$]

This oxide is found in certain native arsenates. It was discovered by Scheele.

Preparation.—By dissolving either As or As_2O_3 in a strong solution of HNO_3; red fumes are given off, As_2O_5 being formed.

When As is fused with potassium nitrate (KNO_3), a salt, potassium arsenate, is formed.

Properties.—A white, fusible, heavy solid, soluble in water, from which solution it may be crystallised as the normal trihydric arsenate (H_3AsO_4).

The arsenates are an important class of salts, presenting great analogy to the tribasic *phosphates*; thus there are arsenates with the following composition:—

$$\begin{aligned} Na'_3(AsO_4)''' &+ 12H_2O \\ H'Na'_2(AsO_4)''' &+ 12H_2O \\ H'_2Na'(AsO_4)''' &+ H_2O \\ (H_4N)'Mg''(AsO_4)''' &+ 6H_2O \end{aligned}$$

As Mg and P may be estimated by means of the insoluble crystalline salt $(H_4N)MgPO_4$ (see p. 90), so can either Mg or As be quantitatively estimated by means of the salt $(H_4N)MgAsO_4$.

The As_2O_5 is less poisonous than As_2O_3; it is reduced to the state of As_2O_3 by means of dihydric sulphite (H_2SO_3); thus:—

$$H_3AsO_4 + H_2SO_3 = H_3AsO_3 + H_2SO_4$$

This reaction is of great importance in analysis, as H_2S only precipitates a yellow sulphide (As_2S_5) very slowly from As_2O_5; the whole of the As may however be rapidly precipitated as sulphide (As_2S_3) from As_2O_3. With silver nitrate, any soluble arsenate forms a *red-brown* precipitate of *arsenate of silver* ($Ag_3.AsO_4$).

Arsenicum with Sulphur.

Arsenicum forms three important compounds with S, two of which are found native, namely, *realgar* and *orpiment*:—

Arsenic sulphide . .	$As''_2 S''_2$	Realgar
Arsenic sesquisulphide .	$As'''_2 S''_3$	Orpiment
Arsenic persulphide . .	$As^{v}_2 S''_5$	

Realgar may be produced artificially, and is used for making *white Indian fire.*

Orpiment ($As'''_2O''_3$) is the sulphide precipitated whenever H_2S is passed through a solution containing As''' and free HCl; *if there be no free* H *salt present, the solution is only coloured yellow, without a precipitate being formed.* It is insoluble in water and dilute HCl or H_2SO_4; it is decomposed by HNO_3. *It is dissolved by ammonia* and the fixed alkalies; also by ammonium sesquicarbonate. It is also soluble in *hydro-ammonium sulphide.* It is a component of the paint *king's yellow.*

ARSENIC PERSULPHIDE.

[$As_2S_5 = 310.$]

THIS salt, sometimes called *pentasulphide of arsenic*, corresponds to the peroxide As_2O_5, and is noted on account of its power to form double S salts. Thus, if through a solution of sodium arsenate (Na_2HAsO_4), the gas H_2S be passed so long as it is dissolved, the O of the salt is removed and S substituted for it; thus:—

$$Na_2H.AsO_4 + 4H_2S = Na_2H.AsS_4 + 4H_2O$$

When HCl is added to this salt, it is decomposed, bright yellow persulphide As_2S_5 being precipitated.

ARSENICUM WITH CHLORINE.

THERE is only one chloride of arsenicum known, the trichloride ($As'''Cl'_3$) or (As_2Cl_6) corresponding to the oxide $As'''_2O''_3$.

The chloride corresponding to As_2O_5 has never yet been obtained.

THE TRICHLORIDE.

[$As_3Cl.$]

MAY be easily formed by bringing metallic As and Cl together. It is a heavy oily-looking colourless fuming liquid, which decomposes when thrown into water.

TRIHYDRIDE OF ARSENICUM OR ARSENIURETTED HYDROGEN.

[H_3As.]

Is a gas of great importance as a test for arsenic; it corresponds to the trihydrides of antimony and of phosphorus.

The gas is produced whenever Zn, H_2SO_4, and a salt of As are mixed together. Like the antimony compound, it burns with flame, and if the tube through which it is passed be heated at one place, the gas is decomposed and a mirror of metallic arsenic deposited inside the tube near that place. If a stream of H_2S be now passed through the same tube when cold, the As is converted into yellow As_2S_3. If now a stream of dry HCl gas be passed through the tube, the yellow sulphide is not acted upon. But if it were antimony sulphide it would be. *This reaction is one means of distinguishing* Sb *from* As. Spots of metallic As are also deposited on a cold body, as a piece of porcelain, for example, if held in the flame of the gas. These spots are of a lighter colour than those produced by Sb under similar circumstances. (See p. 144.)

TESTS FOR ARSENIC.

THE metal is volatile, and gives a garlic-like smell when heated before the blowpipe flame.

As_2O_3 is volatile, and the vapour free from smell.

H_2S produces a *bright yellow* precipitate in salts of As''' and As^v, in presence of free HCl (see p. 150); the precipitate with As^v is but slowly produced, so that the As^v is usually reduced to As''' by means of SO_2 and then tested by H_2S. (See p. 149.)

Copper sulphate produces a green precipitate, called Scheele's green. (See p. 148.)

Silver nitrate gives a *yellow* precipitate with salts of

As''' (see p. 148), and a *red*-brown with salts of As^{v} (see p. 149). For delicate testing an ammonio-sulphate of copper and an ammonio-nitrate of silver are used to test for arsenic. They are formed by precipitating either normal salt with solution of ammonia, and adding ammonia until the precipitate is *nearly* dissolved. The clear solution is used as a test. The solution of As to be tested must be free from any excess of H salt.

Marsh's Test.—By arseniuretted hydrogen (see p. 151, and also antimoniuretted hydrogen, p. 144).

Reinsch's Test.—When bright Cu is boiled with free HCl, the copper remains bright, the HCl having no action on it (see p. 130). But if a solution containing As be boiled with Cu as above, the As is precipitated in the form of a dark grey film of metal on the bright Cu. If such coated copper be carefully dried and heated in a narrow glass tube open at both ends and held in a slanting position, the As oxidises to As_2O_3, and condenses, forming a ring of white octahedral *crystals* in the cold part of the tube.

A similar action takes place with Sb, but the oxide of Sb formed is less volatile, and *amorphous*. (See p. 145.)

GROUP D.—NOBLE METALS.

FIVE metals will be treated of in this group, namely:—

Mercury	Hg	Platinum	Pt
Silver	Ag	Palladium	Pd
Gold	Au		

There are also four other metals generally associated together in the ores of platinum; they belong to this group, but are neither plentiful nor very useful, namely:—

Rhodium	Iridium
Osmium	Ruthenium

There is little family relationship between the metals of this group beyond the fact that in the air they do not tarnish or oxidise; the oxides of the first five are reduced to the metallic state by heat alone, on which account they have received the name of *noble* metals. They do not decompose water under any circumstances, and possess a more powerful affinity for S and Cl than for O.

MERCURY.

[*Symbol*, Hg (Hydrargyrum); *Atomic Proportion*, 200; *S.g.* 13·6.]

Occurrence.—This metal is found in only few localities. It occurs native, but mostly as red sulphide, called *cinnabar* (HgS); occasionally in union with the non-metal selenium (Se), and very rarely as native chloride (HgCl). The chief mines of Hg are at Almaden in Spain and Idria in Transylvania; it is also found in California, Peru, China, and Japan.

History.—It has been known from very remote ages; the alchemists gave it the name mercury (☿); also it is called *quicksilver*, or live silver; its name *hydrargyrum* is derived from *ὕδωρ*, water, and *ἄργυρος*, silver.

Preparation.—Generally from the native sulphide.

1. By heating it in an iron retort with Fe or CaO, when the Hg volatilises and is condensed.

2. By roasting *cinnabar* (HgS) on a grating over a fire; SO_2 escapes, and the fumes of Hg are condensed, either in long tubes or chambers.

Properties.—It is the only metal which is liquid at the ordinary temperature; it is silver-white, and slowly evaporates even without being heated. The vapour of Hg in large quantity is poisonous; care must be taken not to spill the metal in dwelling-rooms or breathe its vapours during experiments. It freezes at -39°C., and is then malleable. It boils at 350° C. Its vapour is colourless. When pure, Hg runs in globules, and does not adhere to glass, porcelain, or paper; but if it contains other metals, it leaves a tail or streak behind it. When pure, it does not tarnish in the air.

Hydric chloride has no action on Hg.

Dihydric sulphate has no action when cold, but boiling hot it attacks Hg with formation of white insoluble mercury sulphate and SO_2, which escapes (see *C. N.* p. 83).

Hydric nitrate, even when cold, attacks Hg, forming a soluble nitrate.

Uses.—Mercury is used to extract gold and silver from their ores by forming an amalgam with them, which is soluble in an excess of Hg. It is used to fill barometer and thermometer tubes; alloyed with tin, to cover the backs of looking-glasses; to make detonating powder, and the colour *vermilion.* Also Hg and its salts are employed as medicines.

MERCURY WITH OXYGEN.

MERCURY forms two well-known compounds with O, representing two classes of salts, which the student must be careful not to confound together, the one containing univalent mercury (Hg'), the other divalent mercury (Hg''), namely—

Black Mercurous oxide or Dimercury oxide .	Hg'_2O.
Red Mercuric oxide or Mercury oxide . .	$Hg''O$.

The black oxide (Hg'_2O) is a powerful basylous radical, and may be formed by adding a solution of alkali to mercury chloride ($Hg'Cl'$).

The red oxide ($Hg''O''$) is sold as a bright red crystalline powder, made by heating Hg, exposed to the air, to nearly its boiling point; O is slowly absorbed, a red powder (HgO) forming on the surface. At a little higher heat, it is decomposed into Hg and O (see *C. N.* p. 50). It may also be formed by decomposing the corresponding *nitrate*

$2(Hg''2NO_3).H_2O$ by heat. It is also precipitated as a yellow amorphous solid by adding a solution of alkali to one of Hg'', as the dichloride ($Hg''Cl'_2$). It is slightly soluble in water.

THE NITRATES OF MERCURY

ARE numerous, the exact composition of the salt formed depending upon the strength of the solution of hydric nitrate used to act upon the metal. It will be sufficient here to state that when Hg is acted upon by *cold dilute* solution of HNO_3 a nitrate is formed containing Hg'; and when a *strong* solution of HNO_3 and *heat* are employed, a nitrate results containing Hg''.

NORMAL MERCUROUS OR MERCURY NITRATE.

$[2(Hg'(NO_3)').H_2O.]$

THIS salt crystallises out in colourless soluble crystals, when Hg is acted upon by dilute *colld* HNO_3. It decomposes when largely diluted with water into an oxysalt.

MERCURIC NITRATE OR MERCURY DINITRATE.

$[2(Hg''2(NO_3)').H_2O.]$

THIS nitrate may be formed by dissolving the red oxide ($Hg''O''$) in HNO_3, or by boiling Hg with strong solution of HNO_3. The normal salt is difficult to crystallise; if largely diluted, a basic insoluble salt is thrown down, from which all the nitrate may be washed by boiling water, leaving $Hg''O$.

MERCURY WITH SULPHUR.

THERE are two sulphides of Hg, corresponding to the two oxides, namely—

Dimercury sulphide	Hg'_2S''.
Mercury sulphide	$Hg''S''$.

The former is precipitated, when H_2S is added to a solution containing Hg'.

MERCURIC OR MERCURY SULPHIDE.

$[Hg''S'' = 232.]$

THIS body occurs native as *red cinnabar*; it is also made artificially as the beautiful *red* paint *vermilion*, but when precipitated by H_2S from a solution containing Hg'', it always appears as a *black* body. There does not seem to be any chemical difference between the black and red variety.

The H salts have little action on $Hg''S''$, nor have the alkalies; it is easily decomposed, however, if heated in the air, its S burning off. The extraction of Hg by roasting cinnabar depends upon this fact.

MERCURY WITH CHLORINE.

THE most important salts of Hg are its two chlorides, namely—

Mercurous or Mercury chloride—*Calomel* .	$Hg'Cl'$
Mercuric chloride or Mercury dichloride—*Corrosive sublimate*	$Hg''Cl'_2$

The $Hg'Cl'$ is *insoluble* in water, and extensively used in medicine. It is *poisonous*.

The $Hg''Cl'_2$ is *soluble* in water, and is a most *virulent poison*. It is also a valuable antiseptic.

MERCUROUS OR MERCURY CHLORIDE.—CALOMEL.

$[Hg'Cl' = 235{\cdot}5.]$

PREPARATION.—This salt is easily prepared as follows:—

1. *Triturate* (from *tritura*, a threshing or grinding), an atomic proportion by weight of Hg with an atomic proportion of $HgCl_2$, and sublime the mixture; thus:—

$$Hg + HgCl_2 = 2HgCl.$$

2. Triturate mercury sulphate ($HgSO_4$) with Hg, when Hg_2SO_4 is formed; thus:—

$$HgSO_4 + Hg = Hg_2SO_4.$$

Then triturate Hg_2SO_4 with 2NaCl and sublime the mixture, when 2HgCl is formed; thus:—

$$Hg_2SO_4 + 2NaCl = Na_2SO_4 + 2HgCl$$

The sublimate obtained by either process must be carefully and repeatedly washed with water to dissolve out any traces of dichloride ($HgCl_2$) with which it may be mixed.

This chloride may also be formed by adding a solution of NaCl to the nitrate $2(HgNO_3).H_2O$, or by passing SO_2 through a solution of the dichloride:—

$$2HgCl_2 + 2H_2O + SO_2 = 2HgCl + 2HCl + H_2SO_4.$$

PROPERTIES.—A yellowish-white, amorphous, heavy, volatile, insoluble, and tasteless powder. It may be crystallised, forming clear and transparent crystals.

It is changed to black oxide by the alkalies or by lime-water, forming *black wash*, from which circumstance it is said to be called *calomel* (from *καλός*, beautiful, and *μέλας*, black).

It is turned black by ammonia (see p. 158).

Either solution of Cl, boiling with HCl, with NaCl, or with $H_4N.Cl$ converts it into $HgCl_2$ which is soluble.

Dihydric sulphate has no action on calomel, but hydric nitrate changes it into a mixture of $HgCl_2$ and nitrate, which are soluble in water.

MERCURIC CHLORIDE OR MERCURY DICHLORIDE.

CORROSIVE SUBLIMATE.

$$[Hg''Cl'_2 = 271.]$$

PREPARATION.—This salt is easily obtained as follows:—

1. By dissolving the red oxide HgO in 2HCl.

2. By triturating $HgSO_4$ with 2NaCl and subliming the mixture; disodium sulphate (Na_2SO_4) remaining in the retort:—

$$HgSO_4 + 2NaCl = HgCl_2 + Na_2SO_4$$

Properties.—A solid, colourless, massive, or in snow-white crystalline needles; volatile, soluble in 16 parts of cold water, burning and disagreeable to the taste; very poisonous. Is used to preserve skins, &c. from decay.

Potash and soda precipitate from solution of $HgCl_2$ yellow oxide (HgO).

Ammonia precipitates a white body of curious composition, the so-called *white precipitate* ($Hg.H_2N.Cl$), the body H_2N (*amidogen*) appearing as a compound radical taking the place of Cl.

$$Hg''Cl''_2 + 2H_3N = H_4N.Cl + Hg''.(H_2N)'Cl'.$$

A similar action takes place when ammonia is added to $Hg'Cl$ (calomel) a black precipitate being formed.

$$2Hg'Cl' + 2H_3N = H_4N.Cl + Hg'_2.(H_2N)'Cl'$$

With *white of egg* $HgCl_2$ forms an insoluble compound. *Eggs* are therefore given as an antidote in case of poisoning with this salt.

MERCURIC or MERCURY SULPHATE.

[$HgSO_4 = 296.$]

This salt is formed whenever H_2SO_4 is boiled with Hg, the gas SO_2 being given off. (See *C. N.* p. 83.) The sulphate deposits as a whitish insoluble powder. This salt is decomposed by water, a yellow basic salt being formed; but is soluble in a solution of NaCl. It is used in the preparation of the mercury chlorides.

TESTS FOR MERCURY SALTS.

All the compounds of mercury are volatile. By heating them in a narrow glass tube with sodium carbonate, globules of metallic mercury are obtained.

They are all reduced when heated with *stannous* chloride, metallic Hg being formed.

A strip of metallic copper becomes covered with a grey coating of Hg if dipped in a solution containing that metal.

Place a drop of solution containing Hg on a piece of gold, and touch the gold and solution with a piece of zinc or point of a knife: a spot of Hg is formed on the gold.

Salts containing Hg'.

Potash, soda, or *lime* gives a *black* precipitate of oxide (Hg'_2O).

H_2S or *hydro-ammonium sulphide* gives a *black* precipitate of sulphide Hg'_2S''.

HCl gives a *white* precipitate of calomel ($Hg'Cl$) *which is turned black by ammonia.*

Salts containing Hg''.

Potash, soda, or *lime*, gives a *bright yellow* precipitate HgO.

H_2S gives, if gradually added, first a white precipitate passing through yellow and red into *black. Hydro-ammonium sulphide* also gives a *black* precipitate, which is not dissolved by HNO_3.

HCl *gives no precipitate.*

Potassium iodide gives a brilliant scarlet precipitate, *which is soluble* in either an excess of potassium iodide or of $HgCl_2$.

SILVER.

[*Symbol,* Ag (Argentum); *Atomic Proportion,* 108; *S.g.* 10·5.]

OCCURRENCE.—Silver is widely distributed, being present in small quantities in many metallic ores, especially those of lead (galena) and copper (pyrites). It is frequently found *native*, but usually as *sulphide*, never as oxide, and only occasionally as chloride (*horn silver*).

The mines of Mexico and Peru are the most celebrated sources of Ag. In Europe it is especially obtained from Saxony, Hungary, and Bohemia; there are also mines in Asia, and it is obtained from many lead and copper smelting works. (See p. 123.)

HISTORY.—Has been used as an ornament and for money from the earliest times. On account of its rarity and beauty, called Luna (☽) by the ancients.

PREPARATION.—The ores of silver commonly contain but a small percentage of that metal; the processes by which it is extracted, therefore, at first seem long and tedious, compared to the amount of metal obtained; the great value of the metal, however, compensates for the enormous labour and expense incurred.

1. The metal, when native, may sometimes be picked in lumps from the ore; when distributed in smaller quantities, the crushed ore is agitated in barrels with mercury. The Hg and Ag form a liquid *amalgam* (see p. 154), from which the excess of mercury is strained, a solid amalgam being left. This is distilled, the Hg passes off, and Ag is left.
2. The metal, when found as sulphide, is generally mixed with sulphides of Cu, Fe, Pb, and other metals. For the extraction of Ag from *galena* (PbS), see p. 123.
   a. The sorted and powdered ore is mixed with NaCl, and gently roasted. SO_2 is given off, and AgCl (silver chloride), together with copper chloride and iron chloride, formed; and also sodium, iron, and copper sulphates.
   b. The roasted ore is put into barrels with water and scraps of iron, and shaken for hours. The Fe extracts Cl from Cu and Ag, which two metals become reduced to the metallic state.

c. Hg is now poured into the barrels, and they are again agitated. An amalgam of silver, copper, and mercury in unknown proportions is thus obtained, which is liquid, because an excess of Hg is always added. This amalgam is separated from the ore and water.

d. The liquid amalgam is strained through cloths, to separate the excess of Hg from the solid amalgam dissolved in it. The solid amalgam which was dissolved in the Hg is distilled, when Hg goes off, leaving a mixture of silver and copper behind.

e. The Cu is separated from the Ag by a process called *liquation* (from *liquo*, I melt), which is the art ot melting a fusible *alloy* from a more infusible substance. To this end, a quantity of Pb is added, and the whole carefully heated, when a fusible alloy of Ag and Pb runs away, the Cu being left behind. The Pb is now separated from the Ag, as on p. 123.

In Mexico, where fuel is comparatively scarce, the mixing of the silver ore and the NaCl is accomplished by the trampling of horses. The process of roasting is done away with, a small quantity of ready roasted copper ore being added instead. The whole operation takes a longer time than the process given above; and much Hg is lost in the form of HgCl; otherwise the two processes are very similar.

PROPERTIES.—The whitest and most lustrous of the metals; the best conductor of heat and electricity; it is also *very* malleable and ductile. Melts more easily than Cu, and, when melted, can absorb a large quantity of O gas, which it gives off again while cooling, often suddenly, the liquid being spattered about.

Silver money is an alloy of 7·5 % of copper and 92·5 % of silver. *Plated* silver is made by heating and rolling together plates of Cu and Ag. *Electro-plate* (silver) is made by depositing Ag, by the aid of a galvanic current, on plates of perfectly clean metal, generally on German silver.

Silver is easily attacked by hydric nitrate, a soluble silver nitrate ($AgNO_3$) being formed. It is also acted on by boiling dihydric sulphate, a soluble Ag_2SO_4 being formed; it is not attacked by HCl.

Silver is not oxidised by exposure to the air; but, from its great affinity for S, it tarnishes, by formation of *black* sulphide, formed by traces of H_2S contained in the air. The purer the silver, the more easily it tarnishes, and the reason why electro-plate silver tarnishes easier than standard silver, is because the silver deposited by the battery is pure Ag.

SILVER WITH OXYGEN.

SILVER forms three compounds with O, only one of which is of importance, namely :—

Argentic or Silver oxide . . . Ag'_2O''

This oxide may be precipitated as a brown powder by adding solution of potash to any soluble salt of silver.

This oxide is *freely soluble in solution of ammonia,* with formation of a most dangerous compound, a *fulminating silver.* Care must be taken when working with solutions of silver and ammonia, or an explosive compound may be formed unintentionally.

True fulminating or detonating powders are made by adding alcohol to solutions of nitrate of silver or mercury. They contain a peculiar radical (C_2N_2O), as $Ag_2C_2N_2O_2$ (the fulminate of silver). They are very dangerous to make.

ARGENTIC OR SILVER NITRATE.

[$Ag'NO_3 = 170.$]

THIS salt, melted and cast into moulds, forms the *lunar caustic* or *lapis infernalis* of the surgeon.

The salt is easily formed by dissolving Ag in moderately strong solution of HNO_3. If silver coin is dissolved, the solution is green, from the presence of copper nitrate.

To separate the Cu, the solution is evaporated to dryness and gently fused, when the copper nitrate decomposes to black CuO: the silver nitrate when cold can be dissolved in water.

The salt crystallises from solution in water in clear, square, anhydrous crystals, easily soluble in water. *Organic matter, as dust, &c., quickly decomposes this solution when exposed to sunlight.* It stains the skin black, and is used for marking linen. These black stains can be removed by solution of potassium cyanide, which dissolves them.

The *bath* of the photographer is a solution of this salt in distilled water.

ARGENTIC OR SILVER SULPHATE.

$[Ag_2SO_4.]$

MAY be formed by boiling Ag in H_2SO_4, when SO_2 is given off. The salt is not very soluble in water.

SILVER WITH SULPHUR.

THERE is only one compound of Ag with S, namely:—

ARGENTIC OR SILVER SULPHIDE.

$[Ag_2S = 248.]$

IT is the principal ore of Ag. Silver has a powerful attraction for S, and easily becomes tarnished in the air or in a room where coal-gas is burning, on account of the traces of S present.

It is easily precipitated as a black solid by passing H_2S through a solution of silver nitrate.

The Ag'_2S'' easily fuses in closed vessels; and when roasted in the air is reduced to metallic Ag and SO_2, which escapes. It is decomposed by strong hot solutions of the H salts.

SILVER WITH CHLORINE.

THERE is only one well-known compound of Ag with Cl, namely :—

ARGENTIC OR SILVER CHLORIDE.

[$Ag'Cl' = 143{\cdot}5$.]

WHEN freshly precipitated, it appears in the form of a white solid, insoluble in water, looking very like *curds* from milk. It fuses and forms a solid horn-like body, which when found native is called *horn silver*. The freshly precipitated white chloride rapidly changes to violet and black by exposure to light.

It is produced by adding HCl or NaCl to a solution of Ag; and as it is quite insoluble in water, is used in analysis to precipitate silver from its solutions.

It is insoluble in dilute HNO_3, *but is dissolved by solution of ammonia* (see *C. N.* p. 57). It is also soluble in sodium hyposulphite ($Na_2H_2S_2O_4$ or $Na_2.S_2O_3.H_2O$); but after it has been blackened by light, it is not soluble in hyposulphite. Upon this fact depends the use of hyposulphite by the photographer. It is also soluble in potassium cyanide. It is easily decomposed by *nascent* H, for if a piece of Zn and a little dilute H_2SO_4 be added to AgCl, the salt is reduced to black metallic Ag with formation of HCl. It may also be decomposed by fusion with Na_2CO_3 or boiling with KHO and grape sugar.

IODIDE AND BROMIDE OF SILVER.

[AgI and AgBr.]

MAY be obtained by means analogous to those employed in the formation of AgCl. They are *both yellowish-white*, and turn black when exposed to the light, and are much used in photography. They are both soluble in sodium hyposulphite, but not after they have been changed by light. They are both soluble in ammonia, *but not nearly so readily as the chloride.*

VOLUMETRIC ANALYSIS OF SILVER.

The term *volumetric analysis* means 'analysis by measure,' in contradistinction to 'analysis by weight;' thus, as AgCl is for all practical purposes absolutely insoluble in water, therefore the smallest portion of silver is precipitated from solution by NaCl (see p. 164). The atomic proportion of NaCl is $23 + 35{\cdot}5 = 58{\cdot}5$, the atomic proportion of Ag is 108: therefore 58·5 grains of NaCl will precipitate 108 grains of Ag from a solution of silver nitrate. The 108 grains of Ag will of course be in combination with 35·5 grains of Cl as AgCl, according to the following formula:—

$$NaCl + AgNO_3 = NaNO_3 + AgCl.$$

If a solution of NaCl be made, containing a known quantity of NaCl (standard solution), we can estimate the quantity of silver in a solution by simply seeing how much of our standard solution is required to precipitate it; without the trouble of collecting, washing, drying, burning and weighing the precipitated chloride. This method of analysis is very commonly employed in laboratories and manufactories where a large number of analyses of one kind has to be made. The instrument used to drop the standard solutions from is called a *burette*. It is a wide and long glass tube, drawn to a point at one end, and very carefully measured into a hundred equal parts, each division containing a known quantity of fluid.

TESTS FOR SILVER SALTS.

Before the blowpipe, mixed with a little Na_2CO_3 on charcoal, they are all reduced to the metallic state.

The hydrates of K *and* Na precipitate brown hydrated oxide *insoluble* in excess.

Ammonia precipitates a brown powder easily *soluble* in excess.

The *soluble carbonates* precipitate white silver carbonate, soluble in ammonium carbonate.

H_2S and *hydro-ammonium sulphide* precipitate black sulphide of silver, insoluble in alkalies, or dilute solution of H salts.

Hydric or any soluble chloride precipitates white chloride of silver, insoluble in HNO_3, *but soluble in ammonia*, sodium hyposulphite, and potassium cyanide.

The *three insoluble chlorides*, namely—

Lead chloride . . .	$PbCl_2$
Mercury chloride . .	HgCl
Silver chloride . .	AgCl

are all white and insoluble both in water and dilute H salts. They may, however, be distinguished from one another by *solution of ammonia*, which has *no action* on lead chloride (see p. 128); *turns* mercury chloride *black* (see p. 159), and *dissolves* silver chloride.

Potassium iodide and *bromide* produce yellowish precipitates in solutions of silver salts, with difficulty soluble in ammonia.

Potassium cyanide precipitates white cyanide of silver, easily soluble in excess of potassium cyanide.

GOLD.

[*Symbol*, Au (Aurum); *Atomic Proportion*, 197; *S.g.* 19·3.]

OCCURRENCE.—This metal, in small quantity, is very widely distributed through the mineral world; it occurs almost wholly native. Large quantities of gold have of late years been brought from California and Australia. Gold is moreover found and frequently *worked* in many other countries.

HISTORY.—Gold has been known from the earliest times, and on account of its rarity and beauty always highly prized as an ornament.

PREPARATION.—By *washing* the sand and mud of many rivers, which contain fine particles of gold. On account of its great weight the gold remains at the bottom of the vessel in which it is washed, while the mud, sand, &c., is carried away by the water.

By *amalgamation*. The ore is crushed and shaken with Hg, similarly as by the working of Ag.

PROPERTIES.—A very lustrous metal of a rich yellow colour, known as *golden*. Not very hard, the most malleable of the metals, and very ductile (see pp. 3, 4). More difficult to melt than copper.

Gold coin is an alloy of Cu and Au; the former is added to give it hardness.

Gold is not acted on by the H salts nor by the fused hydrated alkalies, but is dissolved by a mixture of HCl and HNO_3, called 'Aqua Regia,' because gold has always been considered the king of metals.

GOLD WITH OXYGEN.

THERE are two oxides of gold, Au'_2O and Au'''_2O_3.

GOLD WITH SULPHUR.

A BLACK sulphide of gold Au_4S_4 or $Au_2S.Au_2S_3$ is precipitated when H_2S is added to a solution of gold; it is soluble in sulphides of the alkali metals.

GOLD WITH CHLORINE.

GOLD forms two chlorides, the one AuCl and the other $AuCl_3$.

AURIC OR GOLD TERCHLORIDE.

$[AuCl_3 = 303{\cdot}5.]$

WHEN gold is dissolved in a mixture of three parts of strong solution of HCl and one part of strong solution of

HNO_3, a yellow solution of terchloride is obtained. This, when carefully evaporated to dryness, appears as a deliquescent mass of reddish crystals. The crystals are soluble in water, but unless the chloride contains free HCl it is most easily reduced to metallic gold or to an insoluble Au′Cl′.

A solution of terchloride is the one always used to test the presence of gold. It is very easily decomposed by iron sulphate, antimony terchloride, mercurous nitrate, hydric oxalate, the skin, paper, &c.

Gold purple or purple of Cassius is used to impart to Bohemian glass that rich ruby colour so much admired. It is formed whenever dilute solutions of $AuCl_3$ and $SnCl_2$ are mixed together.

Gold cyanide may be precipitated by adding solution of KCy to dilute solution of gold chloride. It is soluble in excess of KCy.

TESTS FOR GOLD SALTS.

THE solution of gold to be tested should be free from hydric nitrate; then:—

Iron sulphate precipitates *brown* metallic gold;
Tin chloride precipitates *purple of Cassius*;
Mercurous nitrate precipitates brown metallic gold.

All the salts of gold are decomposed when heated in the air.

PLATINUM.

[*Symbol*, Pt; *Atomic Proportion*, 198; *S.g.* 21·5.]

OCCURRENCE.—Only in few places. Native, in flattened grains, commonly associated with palladium, rhodium, osmium, ruthenium, and iridium; some of the crystals which these alloys form are extremely hard. They occur in the Ural mountains, Mexico, St. Domingo, &c.

History.—This metal was first brought into notice by a Mr. Charles Wood, an assayer in Jamaica, about 1740.

Preparation.—1. The ore, which contains about 20% of impurity, is heated in 'aqua regia,' with formation of a red solution of Pt (see p. 170). The red solution is mixed with ammonium chloride, when most of the Pt is precipitated as a yellow nearly insoluble salt of $2(H_4NCl).PtCl_4$. This yellow precipitate is decomposed by heat, leaving metallic Pt as a *spongy* mass behind. It is then pressed into a mould, heated white hot, and welded into a uniform solid mass; or it is melted in the oxyhydrogen gas flame.

2. The platinum ore is roasted in a reverberatory furnace with galena (PbS), and afterwards melted with an additional quantity of litharge (PbO). A fusible alloy of lead and platinum is thus obtained, from which the lead may be removed by oxidation.

Properties.—A grey-white metal, as hard as copper; malleable and very ductile. It cannot be melted in any ordinary furnace, is not dissolved by the H salts, but is attacked by the caustic alkalies; it does not tarnish in the air, and expands less when heated than any other metal.

It is chiefly used in making crucibles and dishes for use in the laboratory.

Spongy platinum possesses the peculiar property of causing a mixture of H and O gases to unite chemically and form H_2O. Perfectly clean plates of platinum also have the power, under certain circumstances, to cause a slow combustion of coal gas.

Platinum black is Pt in a very fine state of division. It is formed when a solution of Pt is heated with sugar and Na_2CO_3. This powder absorbs many times its volume of O gas, and, if ether be dropped upon it, the ether is oxidised in a peculiar manner. It is supposed that Pt, even when in sheets, has the power to condense gases on its surface, and thus brings about chemical union, without itself being in any way acted upon.

PLATINUM SALTS.

PLATINUM forms two oxides, $Pt''O''$ and $Pt^{iv}O''_2$; two sulphides corresponding to the oxides, namely, PtS and PtS_2; the latter may be precipitated by H_2S from Pt solutions, and is slightly soluble in alkaline sulphides; also two chlorides analogous to the two oxides and sulphides, the $PtCl_2$ and $PtCl_4$. There is an immense number of complicated salts of Pt with ammonia and also with cyanogen.

PLATINIC CHLORIDE OR PLATINUM TETRACHLORIDE.

$[PtCl_4 = 340.]$

THIS salt is formed by boiling Pt in 'aqua regia,' and evaporating the solution to dryness at 100° C. It is deliquescent and soluble in alcohol and ether. With other chlorides $PtCl_4$ forms a number of double salts, some of which are nearly insoluble in water. Of these the most important are the *double chloride of platinum and potassium* ($2KCl.PtCl_4$) and the *double chloride of platinum and ammonium* ($2(H_4NCl).PtCl_4$); they are both yellow and quite insoluble in a mixture of alcohol and ether. The former is used as a means of precipitating K from its solutions (see p. 53), and the latter ammonium (see p. 70).

TESTS FOR PLATINUM SALTS.

THE tetrachloride is the only solution of common occurrence; its salts are all decomposed by heat, leaving metallic Pt.

Dilute solution of KHO and dilute solution of ammonia precipitate yellow insoluble double salts.

Hydrate of sodium precipitates a brown oxide.

PALLADIUM.

[*Symbol,* Pd; *Atomic Proportion,* 53·2; *S.g.* 6·8.]

SMALL quantities of this metal are found in the platinum ores. It is remarkable for its power to absorb H gas and to unite with C. When first discovered it caused a great controversy, for it was asserted to be a compound and not an element. It is a silver-white metal, attacked by HNO_3, and forms an iodide nearly insoluble in water.

EXAMINATION QUESTIONS

From Cambridge, London University, and Woolwich Papers.

1. How, and by whom, was potassium discovered? Describe its chief properties, and those of any two of its compounds.

2. When are two elements said to be isomorphous?

3. What is a basyle? Give instances of simple and compound basyles.

4. What is the composition of ferridcyanogen; and how does it differ from yellow prussiate of potash?

5. What is the formula for prussian blue? How is it manufactured?

6. What chemical changes take place in the explosion of gunpowder?

7. What is the property of the alloy of antimony and lead that renders it so valuable for type metal?

8. Describe the processes and decompositions in the manufacture of sodium carbonate from common salt.

9. Give the distinguishing tests for lead, mercury, and silver.

10. How does chlorine act upon a hot solution of caustic potash?

11. What is the effect of heat upon potassium chlorate, if the temperature be gradually raised?

12. Represent, in symbols, the effect of the salts, hydric chloride, hydric nitrate, and dihydric sulphate, upon copper and copper oxide, respectively: first when cold, secondly when heated.

13. An orange-yellow, or yellow precipitate, is produced by transmitting dihydric sulphide through a solution containing free hydrogen salt. What metals may it contain?

14. How is aluminium extracted from clay? Describe the chief physical and chemical properties of the metal.

15. How is antimony obtained from the native sulphide? Give the formulæ for the chlorides, oxides and sulphides of antimony.

16. What is the crystalline form of alum? Express the composition of the group of alums. Name the metallic oxides which, in combination with dihydric sulphate and potash, will form an alum.

17. What is meant by the term compound radical?

18. Represent, in symbols, the changes which occur when sal ammoniac is heated with chalk.

19. Describe a method by means of which iron sesquioxide, black copper oxide, and silver oxide, respectively, may be reduced to the metallic state.

20. How do cast iron, steel, and wrought iron differ from each other in composition?

21. How is lead acted upon by soft water and hard water, respectively?

22. State how chloride of silver may be reduced to the metallic state.

23. Under what condition is arseniuretted hydrogen formed? Describe its leading chemical properties.

24. How is mercury obtained from vermilion?

25. How would you detect the presence of potash, soda, or ammonia in a solution?

26. Describe, by equations, the action of hydric chloride upon antimony sulphide.

27. How do you distinguish the salts of barium from those of calcium?

28. What compounds does mercury form with oxygen and with chlorine?

29. How are the chlorides of mercury made? Specify their principal characters.

30. Describe the oxides and sulphides of arsenic.

31. How is metallic platinum prepared from its ore?

32. Describe the process of the artificial formation of nitre.

33. How may chloride or nitrate of barium be obtained from the sulphate?

34. Describe the preparation of the chlorides of iron.

35. What alloys does copper form with zinc and with tin?

36. A black precipitate has been formed by dihydric sulphide in a solution containing free hydrogen salt. What metals may the precipitate contain?

37. Describe the preparation of potassium iodide.

38. What are the principal ores of zinc? Describe the mode of reducing them to the metallic state.

39. What is the effect of heating the common phosphate of soda to redness?

40. Given a substance which contained any of the following metals, viz. iron, zinc, aluminium, calcium, and magnesium. How would you proceed to determine which it contained?

41. Given a solution containing lime, magnesia, and alumina. How would you demonstrate the presence of these substances?

42. Given some iron sulphate, lead nitrate, and calcium chloride. By what test could you detect the basylous and non-basylous radical in these salts?

43. A mixture of potassium and calcium chloride being given for analysis: demonstrate the presence of the several ingredients, and the absence of other salts.

44. A solution contains a mixture of zinc, iron, and magnesium sulphates. By what test could you show the presence of each ingredient?

45. By what test may magnesia, lime, and baryta be distinguished from one another?

46. What is the action of dilute dihydric sulphate on iron sulphide?

47. Express, in an equation, or a diagram, the results obtained on heating copper turnings with concentrated dihydric sulphate. Name the products, and describe the properties of the gas obtained. If charcoal were substituted for copper, the same gas would be produced, in admixture with another. Give the name, composition, principal properties, and most important sources of the second gas.

48. Some chloride of silver ($AgCl$) is reduced to the state of metal. Two hundred grains of silver are obtained, which are afterwards converted into nitrate of silver ($AgNO_3$). State what quantity of the salt is produced, and how much chloride of silver was originally employed; and give an outline of the methods by which the metallic silver and the nitrate were obtained.

49. An alloy of copper, zinc, and silver being given for analysis, what solvent would you use to dissolve it, and why? Express the reactions by equations.

50. State the chemical action of hydric chloride on iron, rust of iron, silver nitrate, and ammonium carbonate.

51. Distinguish ammonium nitrate and sal ammoniac from the corresponding compounds of potassium.

52. State the action of dihydric sulphide on corrosive sublimate, chlorine, and white arsenic, respectively.

53. If a mixture were given you, which might contain lead, silver, copper, and tin, but no other metals; how would you proceed to analyse it qualitatively?

54. How may caustic potash be distinguished from caustic soda?

---

1. Explain the action of potassium upon water. What are the elements of iron rust? What reaction takes place when the vapour of water is conducted over red-hot iron?

2. Binoxide of manganese (MnO_2) parts with one-third of its oxygen at a red heat. What quantity of the binoxide must be employed to furnish 250 cubic inches of oxygen, and what is the name and composition of the residual substance?

3. A piece of iron is allowed to remain in a solution of 100 grains of sulphate of copper ($CuSO_4$), until the whole of the copper is precipitated. To the solution of sulphate of iron ($FeSO_4$) produced, sufficient caustic potash (KHO) is added to precipitate the oxide of iron (FeO). The latter is separated from the solution of sulphate of potassa (K_2SO_4), and subsequently converted into sesquioxide of iron (Fe_2O_3), by digestion with HNO_3 and ignition.

It is required to know the quantities of metallic copper, of sulphate of iron, of sulphate of potassium, and of sesquioxide of iron, produced in the above experiment.

4. When *sulphate of lead* ($PbSO_4$) is fused with an intimate mixture of *sodium carbonate* ($NaCO_3$) and *charcoal* (C), metallic lead is obtained. Explain by symbols in an equation or a diagram the precise nature of the changes which take place in this reaction.

5. Explain and demonstrate by means of symbols the difference between the action of *hydric nitrate* as a solvent, upon *cuprous oxide*, *cupric oxide*, and upon *metallic copper*.

6. Under what circumstances is water decomposed by the metals potassium, sodium, and iron; and what are the phenomena and results attending the decomposition in each case?

7. What changes occur when a current of air is passed over *galena* (sulphide of lead) heated to redness, and what steps must be afterwards taken with the altered mineral to obtain metallic lead?

8. Small quantities of the following substances: *vermilion* (sulphide of mercury), *orpiment* (sulphide of arsenic), and *sulphide of copper*, are placed respectively, in moderately wide glass test-tubes and exposed to heat for a short time. Describe and explain the results obtained in each experiment.

9. 500 grains of nitrate of copper [$Cu(NO_3)_2$], are dissolved in water, and the solution is divided equally; to one-half a slight excess of solution of potassa (KHO) is added, the other half is evaporated to dryness, and the resulting residue is strongly heated. What are the weights of the *anhydrous* and *hydrated* oxide of copper obtained?

10. Ten grains of bismuth are precipitated from their solution in hydric nitrate by means of dihydric sulphide (H_2S); the substance obtained contains the bismuth united as Bi_2S_3. How many grains of sulphur should this precipitate contain, and how many cubic inches of gas are required for its production?

11. An alloy of silver and copper is dissolved in hydric nitrate; to the liquid is added hydric chloride until no further precipitate is formed. The chloride of silver thus produced, is filtered off and purified by washing. Metallic silver may be obtained from it, either by fusion with sodium carbonate, or by placing it in contact, while moist, with fragments of metallic iron or zinc.

The solution separated from the silver precipitate is made to furnish oxide of copper by being heated and mixed with a slight excess of solution of KHO. The oxide is divided into two parts; the one is reduced to the state of metal by being heated in a current of hydrogen, the other by fusion with potassium cyanide.

Explain, in symbols, the precise nature of the most important steps in the above experiment.

12. Represent in symbols, by equations or diagrams:—

(1.) The decomposition of water by potassium.

(2.) The production of hydrogen from zinc and dihydric sulphate (H_2SO_4).

(3.) The action of hydric chloride upon iron.

Give the name and atomic proportions of each product, and mention to what class of bodies each belongs.

Symbols and Equivalents required: Zn 65, K (*potassium*) 39, Fe (*iron*) 56, Cl 35·5, O 16, S 32.

13. It is required to prepare, for an exhibition of the lime light, two cubic feet of hydrogen. What quantities of dihydric sulphate and zinc will be required for this purpose? What quantity of potassium chlorate ($KClO_3$) is required by theory, to furnish sufficient oxygen for employment with the above amount of hydrogen?

The equivalents required are given above.

100 cubic inches of oxygen weigh 34·2 grains } at 60° Fahrenheit
100 ,, ,, ,, hydrogen ,, 2·14 ,, } 30 inches barometer.

14. I wish to prepare 100 grains of finely divided copper, and have provided myself with a sufficient quantity of blue vitriol (*sulphate of copper*, $CuSO_4.H_2O+4H_2O$). How must I proceed to obtain it from that salt? What becomes of the other constituents of the latter; and what quantity of the salt will furnish me with the above amount of copper?

15. Interpret the changes represented by the following equations:—

$$Sb_2S_3 + 6HCl = Sb_2Cl_6 + 3H_2S.$$
$$Cu_3 + 8(H.NO_3) = 3Cu(NO_3)_2 + 4H_2O + 2NO_2.$$

16. Draw a comparison between the properties of the compounds formed by burning phosphorus, carbon, zinc, sodium, and calcium in oxygen, giving the names and composition of each product.

17. Some galena (sulphide of lead), PbS, *Equiv.* 239, is heated moderately in a current of air for some time. The product of this treatment is afterwards mixed with powdered coal, and the mixture is exposed to a red heat in a crucible. Give a detailed account of the various results obtained in these experiments, and state what quantity of the most important product should be obtained, theoretically, from 1,200 grains of galena.

18. A quantity of copper-filings is divided into three parts. One part is mixed with flowers of sulphur; a second with some tin-filings, and the third with powdered glass. Describe the result of the action of heat upon each mixture.

19. State the chemical names and composition of the following substances: black lead, vermilion, borax, sal ammoniac, laughing gas, blue vitriol, muriatic acid, mosaic gold, green vitriol, saltpetre, white vitriol.

Give a brief account of the characteristic properties of the seven metals included in the composition of the above substances, and state also what non-metallic elements are known, in addition to the seven contained in those substances.

20. Describe and explain the methods of proceeding for obtaining, from the oxide of silver, the oxide of copper and the sulphide of lead, their respective metals in the elementary form.

21. State the chemical nature of the following substances, and the changes effected in the common modes of producing them: plaster of Paris, red lead, chloride of lime, glass, soap.

22. The equivalent of lead is 207; that of oxygen 16; that of carbon 12. State distinctly what these numbers represent, and how they have been arrived at. Give the equivalents of oxide of lead and carbonate of lead, and express their composition in symbols.

23. Some iron filings are mixed with flowers of sulphur, and heat is applied to one portion in a suitable vessel; what will be the result, and by what experiments can you demonstrate the difference in the condition of the iron in the two portions?

24. Give a brief account of the composition and important properties of the substances obtained by burning sulphur, sodium, carbon, zinc, and phosphorus, in air or oxygen; and state what quantity of each of those bodies would be required to furnish one hundred grains of its product of composition.

25. Describe and explain the method of preparing hydric chloride. State the action of that salt upon lime, nitrate of silver, and alumina respectively. If 1 lb. of zinc and 1 lb. of iron be dissolved separately in hydric chloride, which will evolve the larger quantity of hydrogen? Why so?

26. State the order of facility with which the metals copper, sodium, zinc, calcium, platinum, iron, will decompose water; and give the particular conditions under which its decomposition will be effected by each.

27. 500 grains of zinc are dissolved in dihydric sulphate. The gas which is evolved during the solution of the metal is passed through a heated tube, containing 5,000 grains of black oxide of copper. State the name, composition, and quantity, of the substance produced by the action of the salt upon the zinc. What is the gas evolved? What are its characteristic properties? What occurs when it is passed over the oxide of copper, and to what

extent will the weight of the latter be affected by the experiment?

28. Cast iron, wrought iron, bronze, and steel, are used as materials for cannon. Give a brief statement of the chemical nature of each of those materials.

39. If air be passed over copper heated to redness, what are the results? If steam be substituted for air in the experiments, what occurs? If at the conclusion of either of the above experiments carbonic oxide gas is substituted for the air or steam, what results are obtained?

30. Given an alloy of silver and copper, what substances would you use to dissolve it, and what would be the chemical reaction in the process of solution? How could you demonstrate the presence of each metal in the solution?

LONDON: PRINTED BY
SPOTTISWOODE AND CO., NEW-STREET SQUARE
AND PARLIAMENT STREET

Complete in 5 vols. 8vo. price £7 3*s.*

A DICTIONARY
OF
CHEMISTRY
AND THE ALLIED BRANCHES OF OTHER SCIENCES.

FOUNDED ON THAT OF THE LATE DR. URE.

By HENRY WATTS, B.A. F.R.S.

ASSISTED BY EMINENT CONTRIBUTORS.

London: LONGMANS and CO. Paternoster Row.

From the **Midland Counties Herald,** Sept. 17, 1868.

The science of chemistry of the present generation has been remarkable for the great progress it has made; for the extent of its development; and for the great increase in its application to the practical arts. It is, therefore, not only a most interesting branch of knowledge in an abstract sense; but from its intimate connection with the arts and trade, it is indispensable to their perfection and progress; and manufacturers cannot be considered as having been slow to avail themselves of the assistance derivable from it. The dyer, the colour maker, the calico printer, the bleacher, the candle manufacturer, the soap boiler, the glass maker, the electro-plater, the brewer, the artificial manure manufacturer, and the farmer, among others, have all improved, if not actually created, their arts through the assistance which chemistry has been able to render them. Neither have there been wanting many treatises on the science by its very able teachers, from which a sound knowledge of it might be obtained. These treatises have continued to appear with a profusion which is almost confusing, and would be regrettable if it were not for the circumstance that it is a necessary consequence of improvements in theory, and of increase of facts and principles developed by the labours of the many conscientious and indefatigable labourers in this prolific field of discovery.

From the MIDLAND COUNTIES HERALD—continued.

But of all forms of book the most convenient for reference on any subject is the dictionary; and the *Dictionary of Chemistry* of the late Dr. URE, which has been out of print between thirty and forty years, was no exception to this truth. By its great merit it became the standard work of the time, the text-book of the student, and the guide of the manufacturer. Since then, however, chemistry has undergone great changes, and made enormous progress; and the URE's Dictionary of our young days has been left far behind. It is with the object of producing a work worthy of the present state of scientific chemistry that Mr. WATTS, assisted by able contributors, has formed and carried out the project of publishing his new dictionary, which he modestly states to be founded on that of the late Dr. URE, and which is intended as a companion to the new edition of URE's *Dictionary of Arts, Manufactures, and Mines*, to which, therefore, reference is for the most part made for the details of manufacturing operations; but those branches of chemical manufactures which have come into existence or have received important developments since the publication of that work are described in the dictionary as fully as its limits will allow, and in all cases explanations are given of the principles on which manufacturing processes are conducted, and the chemical changes which they involve. On comparing URE's original Dictionary and WATTS's new edition, the points of resemblance are very few. It is true both are necessarily alphabetically arranged, and some few of the articles are copied *verbatim et literatim*; but having made this acknowledgment, we may regard the two books in other respects as totally different. The one contains 840 pages, the other nearly 5,000; the one is characterised by the chemical philosophy of forty years ago, and which, though fully worthy of all respect, is now partially exploded; the other is *au courant* with the most advanced philosophy of the present day.

The number of elements and compounds, particularly of the latter—comparatively few then—has been increased by thousands, as organic chemistry has unfolded its riches to the persevering spirit of enquiry which has of late years displayed itself; and the results are given in the new work in a clear, systematic, comprehensive, and perspicuous manner. Some of the articles are complete essays on the subjects of which they treat, and are written with a clearness and precision which it is always agreeable to find, though not always observable in the writings of scientific men, who too often, indeed, express themselves with a clumsy ambiguity, and not seldom without a reasonable regard to the exigencies of grammar.

The article on Classification may be mentioned as of much importance and value, particularly to the earnest student, embodying as it does the most recent views of the most distinguished chemists of the present day.

The subject of Atomic Weights and Volumes, by Dr. ODLING, is also admirably elucidated. It need hardly be said that it is one which is most

From the MIDLAND COUNTIES HERALD—continued.

intimately connected with those changes which chemical science has recently undergone, and is, indeed, at present undergoing, and must, therefore, be highly acceptable to the scientific chemist. Dr. ODLING has done as much as any English chemist to introduce the new notation and theories, and has succeeded in raising the reputation of English scientific chemistry. While, however, abstract questions have received able philosophical treatment and exposition, the practical portions of the subject have not by any means been neglected. The articles on Fuel by Dr. PAUL, on Metallurgy and on the Chemistry of Geology, by the same author; that on Gunpowder by the editor; also that on copper, in which the results obtained by ABEL and FIELD are embodied; with many others which it is unnecessary to particularise, all show that the work has for its object a combination of the theoretical with the practical. Neither have physics been neglected. No work on chemistry could be considered complete without the embodiment of the principles of heat, light, and electricity, inseparably connected as they are with chemical action; and all of these have received full and able treatment. Crystallography, a confessedly difficult and unpalatable subject, is brought forward in an excellent article by the editor, who has been, also, a large contributor of other valuable articles.

A most excellent feature of the book is the large number of bibliographical references which it contains; on every page they may be found, whereby the opportunity is afforded to the reader for consulting the original memoirs.

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